

SOLVOLYSIS OF ESTERS OF
TRIFLUOROACETIC ACID

CENTRE FOR NEWFOUNDLAND STUDIES

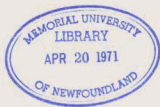
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SOLVOLYSIS OF ESTERS OF
TRIFLUOROACETIC ACID

A Thesis

by

© J. Winter (nee Martin) B.Sc.

Submitted in partial fulfillment
for the degree of Master of Science.

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ABSTRACT

ABSTRACT

Although the spontaneous hydrolysis of simple alkyl esters related to halogenoacetic acids has frequently been studied, no systematic body of rate data is presently available concerning the reactivity of these substances in light and heavy water. The purpose of the present study is to survey the temperature dependence of the rates of hydrolysis in water and deuterium oxide of four esters derived from trifluoroacetic acid (CF_3COOR : R = methyl, ethyl, i-propyl and tertiary butyl).

On the basis of ^{18}O tracer Bunton and Hadwick (1) have concluded that the spontaneous hydrolysis of methyl trifluoroacetate in pure water takes place by acyl-oxygen fission (B_{Ac}^2 mechanism). On the other hand, Moffat and Hunt (2) have concluded that tertiary butyl trifluoroacetate reacts by an $\text{S}_{\text{N}}1$ mechanism in acetone-water mixtures. These observations suggest that a change in mechanism ($\text{B}_{\text{Ac}}^2 \rightarrow \text{S}_{\text{N}}1$) is to be expected in the hydrolysis of the series of trifluoroacetates mentioned above. The aim of this work is to attempt to locate the region of mechanistic change by using the thermodynamic parameters characteristic of the activation process (ΔG^\ddagger , ΔH^\ddagger , ΔS^\ddagger , and ΔC_p^\ddagger) as well as the ratio of the rates of hydrolysis of each substrate in light and heavy water. An attempt has also been made to assess the significance of the parameter, ΔC_p^\ddagger , in cases where a reaction is occurring by more than one mechanism.

INTRODUCTION

INTRODUCTION

The early studies (3) involving carboxylic acids and their esters were carried out in 1862 by Bertholot and St. Gilles who quantitatively examined the reaction of acids with alcohols to produce esters and water. These particular investigations along with similar work by other chemists have a lasting importance in that they led to the formulation of the Law of Mass Action by Guldberg and Waage in 1867. Later Reicher, Van't Hoff, Ostwald and Arrhenius investigated the saponification of methyl acetate at 9.4°C with several alkaline hydroxides. These workers found little difference in rate constants for strong bases such as potassium, sodium, and calcium hydroxide, but a marked decrease in the rate was evident when a weak base such as ammonium hydroxide was used. Further studies involving the effect of salts upon saponification reactions led Arrhenius to conclude that the active reagent in the alkaline hydrolysis reaction was the OH^- ion.

Such early studies concerning the reactivity and equilibria of carboxylic esters served as a basis for subsequent investigations of the mechanism of ester hydrolysis in acidic, basic, and neutral media.

Warder (4), Reicher (5), Holmberg (6) and Prevost (7) studied alkaline hydrolysis of carboxylic esters. Reicher's work on aliphatic esters was extended by Skrabal (8),

Kindler (9) and Olsson (10) who demonstrated the retarding effect of alkyl substituents and the accelerating effect of such groups as $-Cl$, $-CO_2Me$ & $-COMe$. Further studies in ester hydrolysis in the period of 1936-1938 were carried out by Kindler (11), Ingold (12), Evans (13), Hinshelwood (14) and their respective co-workers.

The extensive data from these studies provided Ingold and co-workers with an adequate basis for the mechanistic designations which are currently in use.

CLASSIFICATION OF ESTER HYDROLYSIS

Ingold (16) has classified the mechanism of an ester hydrolysis according to the position of bond cleavage and the molecularity of the reaction. Acyl-oxygen fission (a) or alkyl-oxygen fission (b) may occur.



(a)



(b)

Depending on the reaction conditions two mechanisms may exist for ester hydrolysis. These are analogous to the two mechanisms of nucleophilic substitution or elimination. Ingold (16) denotes these mechanisms in a similar manner and the notation given below summarizes his classification of these processes. The "basic" mechanisms, including both alkaline and neutral hydrolysis are represented by the letter B, and acidic

mechanisms by the letter A. Whether the reaction involves alkyl or acyl-oxygen fission is designated by the subscript Ac or Al and the numbers 2 and 1 indicate whether the reaction is bi- or unimolecular.

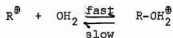
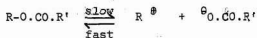
Thus for acid-catalysed hydrolysis and esterification there are two possible mechanisms involving acyloxygen fission, A_{Ac}^1 and A_{Ac}^2 . These two mechanisms cover the more common cases of acid-catalysed hydrolysis and esterification. Both the A_{Ac}^1 and A_{Ac}^2 mechanisms involve a pre-equilibria with the addition of a proton. In the A_{Ac}^1 reaction, heterolytic fission occurs and a carbonium ion is formed, which then reacts with the solvent and loses a proton to form the acid. The two steps of the reaction, after the addition of a proton, are similar to the S_N1 reactions proposed for carbonium ion formation in nucleophilic substitution reactions.

The A_{Ac}^2 mechanism involves a slow reversible step whereby a water molecule attacks as the HOR group leaves the reaction site in a manner similar to S_N2 nucleophilic substitution.

Alternatively acid-catalysed hydrolysis may proceed by alkyl-oxygen fission and with the appropriate choice of the leaving alkyl group the mechanism A_{Al}^1 has been observed.

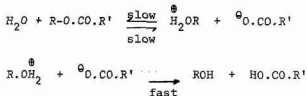
Under certain conditions alkyl-oxygen fission

(B_{Al}¹ and B_{Al}²) will occur in the case of catalysis by alkali or water. In an ester molecule there are two carbon atoms which are susceptible to nucleophilic attack; the carbonyl carbon atom and the α carbon of the alkyl group. However, since the carbonyl carbon is unsaturated, there will be a much stronger tendency towards attack at this position and, as a rule, acyl-oxygen fission is a great deal more common than the alkyl-oxygen fission. It is possible, nevertheless, that both reactions occur and the former dominates but in the case of a slower reaction in an appropriate solvent, the latter becomes more important and consequently observable experimentally. The B_{Al}¹ mechanism may be written in the form (16):



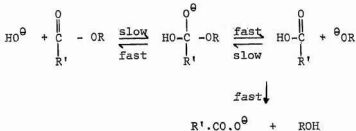
The last step drives the reaction in the direction of hydrolysis. Electropositive groups on R and electronegative groups on R' should accelerate the reaction and there should be no retardation due to steric effects (16).

Mechanism B_{Al}² has been observed in special cases (17) and may be formulated as follows (16):

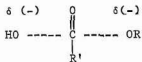


An optically active group (R) should become completely inverted on hydrolysis by the B_{Al}² mechanism.

The most common mechanism of base hydrolysis is the B_{Ac}² mechanism which involves acyl-oxygen fission and follows a second order kinetic law when the attacking reagent is OH[⊖]. Theories suggest that only the bimolecular process takes place and, in fact, the reaction may be represented in the following way (16):



The model of nucleophilic substitution may be used where the reaction intermediate is a complex of the form:



and the attacking and leaving groups share the negative charge.

The rates of hydrolysis of esters by this mechanism are affected by changes in structure. Since the intermediate complex is negatively charged, electropositive substituents would be expected to retard and electronegative substituents to accelerate the reaction. Since the reaction is bimolecular, steric effects near the reaction site should be of great importance in determining the rate of hydrolysis,

Spontaneous $B_{AC}2$ ester hydrolysis occurs in neutral solution at a measurable rate if the R' group of the ester $R'.CO.OR$ is sufficiently electron withdrawing to create a strongly positive centre at the carbonyl carbon where nucleophilic attack by H_2O will occur. The R' group should not be too bulky so as to hinder the approach of a water molecule to the proposed reaction site.

A most favourable instance of this type of mechanism is likely to be provided by esters derived from the halogenoacetic acids. The rates of reaction for a series of halogenated ethyl esters are given in Table I and the plot of $\log k$ vs pK_a in Figure 1 is reasonably linear. This is consistent with the idea that the main driving force for the reaction is the nucleophilic attack by the water molecule on the acyl carbon of the ester, and that changes in the rate constants are

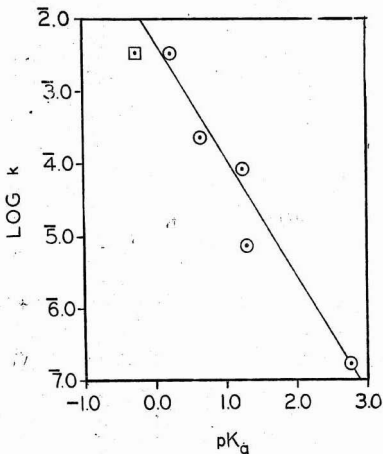


Figure 1

CORRELATION OF RATES OF HYDROLYSIS OF ETHYL HALOGENOACETATES
WITH THE pK_a OF THE CORRESPONDING ACID AT 25°C

TABLE I

RATES OF HYDROLYSIS FOR SOME ETHYL HALOGENOACETATES
IN WATER AT 25°C AND ZERO IONIC STRENGTH

<u>Ester</u>	<u>pK_a of acid</u>	<u>k(sec⁻¹)</u>	<u>log k</u>
CF ₃ COOEt	0.23 (-0.26*)	3.39 × 10 ⁻³	3.53
CCl ₃ COOEt	0.65	2.41 × 10 ^{-4**}	4.38
CHF ₂ COOEt	1.24	8.7 × 10 ^{-5†}	5.94
CHCl ₂ COOEt	1.29	7.7 × 10 ^{-6‡}	6.89
CH ₂ ClCOOEt	2.86	1.7 × 10 ^{-7‡}	7.23

* value determined by Redlich and Hood (18)

** Kurz (19)

† extrapolated from values obtained by Jencks and Carriuolo (20) at ionic strength equal to 1

‡ except where stated from reference (21)

related mainly to the change in the pK_a 's of the acids corresponding to the esters, i.e. to the electron withdrawing ability of the R' group.

Before discussing the driving forces for spontaneous ester hydrolysis, a review of the reactions of esters similar to those used in this particular study is informative.

HYDROLYSIS OF ALIPHATIC ESTERS

A great number of the early studies of ester hydrolysis dealt with aromatic derivatives. However, Taft has discussed the hydrolysis of aliphatic esters and has collected much information dealing with the inductive effects of both acyl and alkyl substituents. He has described the approximate linear free energy relationships between rates of esterification and hydrolysis for esters, which do not follow the Hammett Equation. The relationships, for all the data he collected, were tabulated by relating the logarithm of the relative rate, k/k_0 , (k_0 is the rate constant for the reaction of an ester chosen as a standard) for various reaction series to the same quantity for one series chosen as a standard of reference.

The defining equation was given as (22):

$$\log k/k_0 = f A \quad (1)$$

where f is a proportionality constant dependent upon the

nature of the reaction series; k is the rate constant for any ester in the series; k_o is the rate constant for acetate or ethyl ester in a given aliphatic reaction series or for ortho-substituted benzoates; $A = \log k/k_o$ for the standard reaction series (for which $f = 1$). The quantity 'A' is a substituent constant dependent only upon the nature of the group introduced so as to derive the given ester from the standard ester. Taft (23) has also determined E_o values, which give more or less quantitative measures of the polarities for some aliphatic substituents. These parameters, as well as E_s , which give a measure of the steric factors associated with the substituents, were determined from rates of esterification and hydrolysis of esters.

In later papers (24) (25) Taft and co-workers use these and further rates to discuss the nature of the polar effects of substituents and they then evaluate the resonance effects on reactivity by applying the linear free energy relationships.

Although the hydrolysis of aliphatic esters has been extensively studied by Taft and others, the aliphatic halogenoacetates have been comparatively neglected. However, the solvolysis of halogenoacetates has been studied from a mechanistic standpoint in several instances, usually in mixed solvents in the presence of acids and bases.

THE HYDROLYSIS ESTERS DERIVED FROM THE HALOGENOACETIC ACIDS

The first investigators of these types of compounds were Skrabal and Ruckert (26), who in 1928 determined by a titration method the rate constants at 25°C for the acid, alkaline and neutral hydrolysis of methyl monochloroacetate and dichloroacetate.

Timm and Hinshelwood (27) in 1938 determined the Arrhenius parameters for the acid hydrolysis of ethyl mono-, di-, and tri- chloroacetates from rates determined at ca. 25, 40, 60, and 80°C in alcohol water and acetone water mixtures. They found that E is higher in acid hydrolysis than in alkaline hydrolysis and the order of reactivity of the esters to be tri- > mono- > dichloroacetate from 25-60°C.

In 1939 Palomaa, Salmi and Korte (28) measured the rate of acid hydrolysis of methyl and 2-methoxyethyl trifluoroacetate in water and a 50/50, (V/V), water/dioxan mixture. They determined the temperature coefficients $k_{35}:k_{25}$ and $k_{25}:k_{15}$ for these reactions, and the physical constants for several other trichloroacetates were also recorded. In the following year Salmi and Suonpää (29) reported some investigations concerning the spontaneous water hydrolysis of some chloroacetates. They calculated temperature coefficients for mono-, di-, and tri-chloroacetic acid esters and found the results similar to those obtained in 1939.

In 1941 Anantakrishnan and Krishnamurti (30) attempted a systematic study of a series of haloacetates. They investigated the acid-catalysed hydrolysis of the ethyl esters of monohaloacetic acids (XCH_2COOH ; $\text{X} = \text{Cl}, \text{Br}, \text{I}$) and monohalopropionic acids in 60/40, (W/W), dioxan/water, where the halogens were Cl, Br, and I. Rates were compared to the rates for the unsubstituted acetic and propionic esters and the difference in rate was assumed to arise only from the inductive effect of the halogens. The effect of the halogens on the Arrhenius parameters was discussed from the standpoint of the electronic theory of organic reactions.

In 1950, Bell and Prue (31) studied the kinetics of hydrolysis of methyl monochloroacetate in the heterogeneous buffer system, quinine, quinine sulphate and potassium sulphate at pH 7-8. These workers were mainly interested in the behaviour of several buffers and the choosing of a most effective buffer for a kinetic study of the hydrolysis of several species which are very sensitive to hydroxyl ion concentration. They did find, with the above mixture, that changes in concentration of the buffer component had no effect on the rate of hydrolysis of the methyl monochloroacetate, i.e., there is no heterogeneous catalysis.

Traynam (32) esterified (-) 2-octanol with

trifluoroacetic acid and perfluorobutyric acid and subsequently carried out the aqueous alkaline hydrolysis of the two esters. Because the optical rotation of the alcohol was retained after hydrolysis, he deduced that the reaction took place by acyl-oxygen bond fission which is consistent with the mechanism accepted by Day and Ingold (15) for the formation and hydrolysis of esters of secondary alcohols and other carboxylic acids.

Gorin et al (33) hydrolysed ethyl tri-, di-, and monofluoroacetates in 70% by volume aqueous acetone. In the case of the ethyl monofluoroacetate, the acid catalysed reaction predominated, and both the spontaneous hydrolysis by water and possible autocatalysis by liberated acid were unimportant. The hydrolysis of the difluoroacetate was found to occur at an appreciable rate in initially neutral solution, and autocatalysis due to the difluoroacetic acid released during the reaction was also detected. In acid catalysis, the spontaneous reaction only account for about 10% of the total rate. The hydrolysis of ethyl trifluoroacetate in initially neutral medium was rapid and no autocatalysis by the hydronium ion released in the reaction was evident. However, acid catalysis by 0.05M. hydrochloric acid could be observed. Energies of activation were calculated using the Arrhenius Equation. The mechanism proposed for the catalysed reaction

was again the same as that proposed by Day and Ingold (1) for acid catalysis of esters whereby the reaction is governed by two processes, namely, the extent to which the ester will accept a proton at equilibrium and the rate at which the resultant complex is hydrolysed.

Nair and Amis (34) investigated the acid hydrolysis of ethyl dichloroacetate in several acetone-water mixtures at 25, 35, and 45°C. As Fairclough and Hinshelwood (35) had observed in other cases there was a linear correlation between $\log PZ$ and $1/\sqrt{E}$.

Concurrently, Tommila (36) studied the kinetics of alkaline and acid hydrolysis of ethyl chloroacetate and β -chloroethylacetate. Hydrolysis of the former occurs both in water-acetone mixtures and in water, but the latter only in water. Tommila (36) compared the orders of reactivity in acid and found ethyl formate > acetate > chloroacetate and interpreted the change in rate in terms of the change in A and E in the Arrhenius equation. In alkaline hydrolysis the order of reactivity is chloroacetate > formate >> acetate. These results agree with electronic theory and are in harmony with the results of Gorin, Pierce, and McBee (33) concerning the hydrolysis of ethyl fluoroacetates.

Several papers by Euranto and his co-workers appeared

later (37), (38), (39), (40) and dealt with a system similar to that studied by Tommila (36). Euranto investigated the acid catalysed (38), (39) hydrolysis of a series of α -chloro-alkyl acetates which included α -chloroethyl monochloroacetate in acetone-water mixtures. In the case of the last mentioned ester the mechanism involves a nucleophilic displacement of the α -chloro atom. He also studied the alkaline hydrolysis of a similar series of chloromethyl esters including the chloroacetic acid ester. The data for the series of chloromethyl esters conformed to the Taft equation as well as data from other esters, indicating that the hydrolysis takes place by a B_{AC}^2 ester hydrolysis mechanism and not by nucleophilic substitution of chlorine.

Hornyak and Amis (41) published work in which they discussed the neutral and acid hydrolysis of ethyl dibromoacetate in acetone-water mixtures and the variation in the rate constant with the dielectric constant of the medium.

Moffatt and Hunt (42) published the first of three papers in which they discussed the rates of hydrolysis of fluorinated esters in general and trifluoroacetates in particular. They found that the pseudo first order rate constants for n-alkyl trifluoroacetates can be related quantitatively to the molecular weight of the n-alkyl group and that the pseudo first order rate constants for the ethyl esters of trifluoroacetic, pentafluoropropionic and heptafluorobutyric acids can

be related quantitatively to the molecular weight difference. The reactions were carried out in acetone (70% by volume) -water mixtures and were followed by conductance. The rate constants were calculated from the equation:

$$k = 1/(t_2 - t_1) \ln [(C_\infty - C_1) / (C_\infty - C_2)] \quad (2)$$

where C_∞ is the specific conductance at infinite time, C_1 is specific conductance at time t_1 and C_2 , at time t_2 . Activation energies and the Arrhenius parameters E_a and $\log PZ$ were recorded. The rate constants for the hydrolysis of the trifluoroacetates were related to the structure of the n-alkyl group by the equation:

$$\ln k = a/M + b \quad (3)$$

from which can be derived the equation

$$\log(k/k_0) = a(M_0 - M) / 2.303M_0M \quad (4)$$

where M_0 is the molecular weight of CH_3- , M the molecular weight of R for the compound in question, k_0 is the rate constant for the standard member of the series (the methyl ester), and k is the rate constant for the compound being studied. This equation is closely related to the Taft equation mentioned earlier

$$\log(k/k_0) = fA \quad (5)$$

where f is a constant dependent upon the nature of the series and A is a substituent constant. If f is equal to 1, then $A = a(M_0 - M) / 2.303M_0M$ and this was found to be true in the case of the trifluoroacetates.

In the second paper of the series (2) Moffat and Hunt investigated the hydrolysis of tertiary alkyl trifluoroacetates under similar conditions, and calculated first-order rate constants, relative rates, % elimination, activation energies and activation entropies. The percent elimination varied from 15 to 33%, indicating that the reaction took place via carbonium ion formation (S_N1 - S_N2 mechanism). They proposed that the most likely mechanism is the slow formation of a carbonium ion

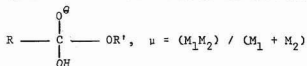


which reacts rapidly with the water and loses a proton to form the acid. This is similar to the mechanism for the hydrolysis of tertiary alkyl halides (43) and a comparison of the data obtained for the two series at 25°C shows a close parallel between them. This also suggests that both types of reaction occur by the same mechanism.

The third paper (44) of the Moffatt and Hunt series again deals with chain length effects in the neutral hydrolysis of some secondary alkyl and aryl trifluoroacetates and in this paper they discuss the large differences that are observed between acid-catalysed and base catalysed acyl-oxygen fission in connection with the chain length effects and solvent effects. Autocatalysis due to liberated acid was evident after 40-50% reaction of the secondary alkyl trifluoroacetates, but no such acid catalysis was observed in the case of the aryl trifluoroacetates since the aryl groups superior electron withdrawing

properties compared to alkyl groups.

The activation parameters (ΔS^\ddagger and ΔH^\ddagger) were tabulated and the effect of solvent (several acetone-water mixtures) on the rate constants for ethyl and isopropyl trifluoroacetates were given. A new equation, relating the rate of hydrolysis and the reduced mass of the intermediate,



where M_1 is the mass of the group to the left of the bond which is broken and M_2 is the mass of the OR' group to the right of the bond, is given in the form

$$\ln k = a/\mu + b \quad (6)$$

Good correlations were obtained for the secondary alkyl trifluoroacetates and benzoates.

In these publications by Moffatt and Hunt, little attempt was made to draw any correlation between the thermodynamic parameters which were calculated and the reaction mechanism, and the reaction was discussed solely in terms of rate constants and molecular weight correlations.

At the same time Bunton and Hadwick (1) were studying ester hydrolysis in dioxan-water mixtures by means of

tracer methods. Polanyi and Szabo (45) originally described this method whereby 18O enriched water could be used to determine whether an ester hydrolysis occurred by an acyl- or alkyl-oxygen bond fission mechanism. Methyl, phenyl, and diphenylmethyl trifluoroacetates were three of the esters used for these investigations by Bunton et al. By studying 18O content of the reaction products it was determined that the acyl-oxygen bond is broken during the hydrolysis in acid, alkaline or initially neutral hydrolysis of methyl and phenyl trifluoroacetates and that the alkaline hydrolysis of diphenylmethyl trifluoroacetate proceeds by a $\text{B}_{\text{Ac}}2$ mechanism.

The Arrhenius parameters derived for the methyl ester were found to be in the range usually associated with bimolecular ester hydrolysis involving acyl-oxygen fission and the marked effect of the addition of alkali to the reaction mixture indicated that the mechanism is $\text{B}_{\text{Ac}}2$. The evidence indicates that probably the slow step of the reaction is nucleophilic attack of a water molecule on the acyl-oxygen carbon atom, i.e. $\text{B}_{\text{Ac}}2$ mechanism. Whether the bond-making and breaking occur simultaneously or an intermediate of finite lifetime is formed is not known.

Autocatalysis due to the trifluoroacetic acid formed in the reaction was observed and the increase in rate was close to that observed in acid-catalysed reaction with the same

initial concentration of acid. Neutral salts were found to decrease the rate of hydrolysis in initially neutral solution and this was explained by Bunton and Hadwick in terms of activity coefficients of the initial and transition states and the activity of water which is decreased with dissolution of neutral salts. The acid-catalysed rate of hydrolysis was slower than the neutral rate and they assumed this was caused by a decrease in the basicity of the neutral ester molecule.

Increasing the concentration of perchloric acid produced a levelling off of the rate constants until a maximum rate was obtained. Bunton and Hadwick asserted that this levelling off is probably not due to the complete protonation of the neutral molecule since a trifluoroacetate would be too weak a base in ca. 1M. perchloric acid in water-dioxan mixtures but arises because the ions of perchloric acid produce a salt effect and decrease the rate of both the acid-catalysed and neutral reactions.

Bunton and Hadwick (1) have assumed that the kinetic form of the acid hydrolysis of the methyl trifluoroacetate is A_{Ac}^2 since the mechanism A_{Ac}^1 is found only in highly ionizing solvents or in the case of esters in which the acylium ion is stabilized by electron release.

The same mechanisms and similar Arrhenius parameters were obtained for the hydrolysis of the phenyl trifluoroacetate. A low value of the non-exponential term in the Arrhenius

equation was equivalent to a negative entropy of activation of about $\Delta S^\ddagger = -40 \text{ e.u.}$, which suggests a great deal more organization of the water molecules in the transition state than in the initial one. This would lead to the conclusion that the transition state is similar to that proposed by Bender (46) for addition to a carbonyl carbon.

The diphenylmethyl carbonium ion is known to be relatively stable and for this reason some alkyl-oxygen fission in the acid and neutral hydrolysis of the corresponding ester was anticipated. Tracer experiments showed that the acyl-oxygen bond is broken during alkaline hydrolysis and that mixed bond cleavage is found for hydrolysis in acid and initially neutral solution. In neutral solution the B_{Al1} mechanism is favoured relative to the B_{Ac2} mechanism by the temperature and solvent composition which favour carbonium ion formation. Many unimolecular processes have higher activation energies than similar bimolecular reactions and consequently the rate of alkyl-oxygen bond fission will increase with temperature. Increasing the water content will facilitate charge separation and therefore increase alkyl-oxygen cleavage as well.

Competition experiments using sodium azide (1) indicated that the carbonium ion formation by the B_{Al1} mechanism was much greater than that shown by tracer experiments but, in fact, the azide probably favours carbonium ion

formation because of an ionic strength effect and at the same time retards hydrolysis by the B_{AC}^2 mechanism through a salt effect.

Bunton et al found that, compared to other ester hydrolysis, the extent of acid-catalysis of the hydrolysis of diphenylmethyl trifluoroacetate is not large but it is greater than for the other trifluoroacetates. In this case, as in the alkaline and neutral hydrolysis, the tracer studies showed that two mechanisms were being followed, A_{Al}^1 and A_{AC}^2 , where A_{Al}^1 has the greater activation energy. It was found that the acid hydrolysis of diphenylmethyl trifluoroacetate fits neither of the usual limiting forms, i.e., the rates of hydrolysis were proportional neither to the acid concentration in bimolecular nor to Hammett's acidity function h_0 in unimolecular reactions. This may be a consequence of two simultaneous reactions of different kinetic forms. However, it was found by these and further isotopic and kinetic experiments that the extent of the unimolecular mechanism, A_{Al}^1 , increases relative to A_{AC}^2 with increasing temperature and perchloric acid concentration.

In 1961 an extensive study of the general base catalysis of ester hydrolysis was carried out by Jencks and Carriuolo (20). They reported the rates of hydrolysis of several esters including ethyl mono-, di-, and trichloroacetates

and ethyl difluoroacetate in the presence of increasing buffer concentrations at a constant pH. They concluded that the general base catalysis of acyl-activated ester hydrolysis seems to represent classical base catalysis rather than nucleophilic catalysis for several reasons:

- (1) The solvent isotope effect k_{H_2O}/k_{D_2O} , has a value from 2 to 3 which is an indication that a bond to hydrogen is stretched in the transition state of the catalysed reaction and that the catalysis involves proton transfer.
- (2) Whereas imidazole is about 4,000 times more reactive than the phosphate dianion as a nucleophilic catalyst in the hydrolysis of p-nitrophenyl acetate (47) (48), the two compounds of almost equal basicity are almost equally effective as catalysts for the hydrolysis of ethyl dichloroacetate. This suggests that the rate depends upon the basicity rather than the nucleophilicity of the catalyst. A Brønsted plot of $\log k_2$ vs pK_a with a slope of 0.47 for the base catalysed hydrolysis of ethyl dichloroacetate strengthens this idea, since the pK_a 's will give a measure of the basicity of the catalysts used.
- (3) With aniline as catalyst, the hydrolysis of ethyl dichloroacetate does not produce dichloroacetanilide as a product. This is another indication that the reaction is not nucleophilic.

In the case of the neutral hydrolysis of ethyl

dichloroacetate Jencks et al (20) found that water can be considered the catalyst in the reaction since it falls on the Brønsted plot. This means that at least two water molecules are required in neutral hydrolysis, one of which is a proton transferring agent. This could account for the large solvent isotope effect.

Jencks and Carriuolo (20) concluded that, in the case of esters with a good leaving group relative to the attacking group, the rate limiting step is sensitive to the nucleophilic reactivity of the attacking reagent, whereas with a relatively poor leaving group, the transition state reflects the expulsion of this group from a species resembling a tetrahedral (46) addition compound. In the case of a poor leaving group the nucleophilic reactivity becomes of secondary importance and the proton transfer, to aid departure of the leaving group and to prevent expulsion of the attacking group, becomes an important part of the reaction.

Recently (1963) Euranto and Cleve (40) have studied the neutral hydrolysis of chloromethyl chloroacetate. They investigated the reaction by an argentometric method at 10° intervals from 5°C to 85°C. Because the calculated activation energy was found to decrease with increasing temperature, the Arrhenius Equation was not considered valid and the following

three parameter equation (49) was used:

$$\log k_O = A/T + B \log T + C \quad (7)$$

and was fitted to the values of k_O for the temperature range from 5° to 65°C by a method of least squares. Thermodynamic parameters, ΔC_p^\ddagger , ΔH^\ddagger , and ΔS^\ddagger , were calculated according to the theory of absolute reaction rates, and the heat capacity of activation was found to be -40 cal mole⁻¹ degree⁻¹ which is close to the value for chloromethyl formate which reacts by the same $B_{AC}2$ mechanism (50). They found that neutral salts and acetone retarded the reaction. Preliminary experiments in heavy water give a solvent isotope effect. k_{H_2O}/k_{D_2O} , of 3.5.

Although a great deal of work has been carried out in the field of ester hydrolysis, it appears that investigations into the solvolysis of haloacetates have not been extensive but nevertheless have provided some insight into the mechanism of neutral ester hydrolysis and acid and base-catalysed hydrolysis.

It appears to be generally accepted (20) (1) that the neutral and base-catalysed hydrolysis of primary and secondary trifluoroacetates take place by a $B_{AC}2$ mechanism. The Arrhenius energies of activation for these hydrolysis reactions are in the same general area of 10 Kcal. mole⁻¹ (42). The tertiary trifluoroacetates have a considerably higher activation energy

in the region 24 Kcal. mole⁻¹ (2) and produce a large percentage of elimination product indicating that the reaction takes place by the formation of a carbonium ion intermediate (S_N2-S_N1 mechanism). The S_N2-S_N1 mechanism is denoted as such to indicate that the reaction does not take place by a typical ester hydrolysis mechanism, e.g., B_{Ac}², but takes place by attack of the nucleophile, in this case water, upon the alkyl carbon of the ester. Whether the reaction is unimolecular or bimolecular is relatively unimportant. The reaction, in fact, probably takes place by a combination of the S_N1 and S_N2 mechanisms although the stability of the tertiary carbonium ion would favour the S_N1 mechanism.

Because of the low solubility of trifluoroacetates in pure water, most of the previous studies have been carried out in solvent mixtures, usually acetone-water or dioxane-water mixtures. However, the initially neutral hydrolysis reactions are easily followed by a conductance method using very low concentrations of ester and with the information derived from earlier work it will be possible to discuss the mechanism of a series of trifluoroacetates.

SOLVOLYSIS IN LIGHT AND HEAVY WATER AND THE SOLVENT ISOTOPE EFFECT

Since the discovery of a second isotope of hydrogen (51) and the isolation of pure deuterium oxide by Lewis and

MacDonald (52) many investigations have been made into the effect of the replacement of hydrogen with deuterium in reactions. In particular, in the period from 1955 to the present time, the rates of hydrolysis in light and heavy water have been determined in the hope of providing further insight into the mechanism of hydrolysis reactions. The interpretation of investigations does not appear, as yet, to provide a satisfactory theory of solvent isotope effects that will be applicable to all hydrolysis reactions.

Swain, Cartinaud and Ketley (53) studied the hydrolysis of some alkyl halides and sulphonium salts in light and heavy water and found the solvent isotope ratio k_{H_2O}/k_{D_2O} to be 1.0 except in the case of tertiary butyl chloride where they found the ratio to be ca. 1.5. They explained this on the basis of the greater electrophilic character of light water, since the tertiary butyl chloride is much more dependent on electrophilic solvation in the rate determining step and is more discriminating between different electrophilic reagents than are the methyl halides. In the later paper, Swain and other workers (54) used the solvent isotope effect to determine the mechanism of ethylene oxide formation from 2-chloroethanol. They deduced that is only secondary effects are operating, a reaction which involves an increase in solvation in going from reactants to transition state should go more slowly in deuterium oxide than in water.

This explanation by Swain and his co-workers was considered to be oversimplified by Laughton and Robertson and in several publications (55), (56), (57), (58), (59), (60) they attempted to interpret the solvent isotope effects in some simple displacement reactions. They studied the solvolysis of some halides, benzenesulphonates, and methyl methanesulphonate in D_2O and H_2O and reported the k_{H_2O}/k_{D_2O} ratios. Unlike Swain et al (53), Laughton and Robertson found that the solvent isotope effect for methyl chloride and methyl bromide was not equal to 1.0 but was in the region of $k_{H_2O}/k_{D_2O} = 1.25$. It was found that the isotope ratios varied with a change in the cation but remained essentially constant with a change in the anion.

Laughton and Robertson (57) discussed the solvent isotope effects from the point of view that water forms a solvation shell of the gas-hydrate type around the substrate in the initial state and, during the activation process, the solvation shell undergoes some sort of reorganization in the vicinity of the reaction site. There are small differences in the work required for this reorganization in light and heavy water and this, they believe, leads to a difference in the rates of solvolysis in the two solvents. This idea that the solvent isotope effect arises from initial state differences is only one interpretation of such effects.

The isotope ratios have been discussed in terms of differences in the solubilities of the transition states rather than of the initial state. Swain and Thornton (54) consider the variation in the ratios are a result of differences in the energy of the transition state only and do not consider the original difference between the free energies of D_2O and H_2O to be important.

Bunton and Shiner (61) have treated solvent isotope effects from a hydrogen bonding point of view and have provided an approximate method for the theoretical estimation of the kinetic isotope effect in terms of the relevant hydrogen stretching frequencies from an assumed model of the transition state. The values obtained by this method are in satisfactory agreement with the experimental values. The treatment is, however, based solely on zero point energies and hence may not be completely valid.

The compounds studied by Robertson and Laughton, and Swain and Thornton have had k_{H_2O}/k_{D_2O} ratios in the region 1.0 - 2.0. Solvent isotope effects have been determined for the spontaneous hydrolysis of some carboxylic esters (62) (63) and anhydrides (20) and kinetic studies in D_2O and H_2O have produced solvent isotope effects which are considerably higher; in the region 2-5. Jencks and Carriuolo (20) found the following values for two haloacetates: ethyl difluoroacetate = 2.1 and

ethyl dichloroacetate = 5, for the hydrolysis of these esters at ionic strength equal to 1.0 in pure water and deuterium oxide.

Bunton and his co-workers (63) have studied the hydrolysis of some carboxylic anhydrides in light and heavy water, and in pure water the solvent isotope ratios are approximately 3. They suggest that the large solvent isotope effect indicates that a change in the structure of the water around the reacting molecules (c.f. Robertson and Laughton) does not account for the entire effect. They suggest a possible mechanism whereby a fast reversible reaction is followed by a rate-limiting process involving a slow proton transfer, but since the effects are large for compounds with no exchangeable hydrogen atoms they suggest that the effect could be explained in terms of the hydrogen bonding between water and the acids and bases concerned (64). On the basis that there is good agreement between predicted and experimental isotope effects, the transition state for the hydrolysis of carboxylic anhydrides is that for the formation of a tetrahedral intermediate (61). This has also been suggested as the intermediate for the spontaneous hydrolysis of some carboxylic esters (46).

There appears to be no definite agreement as to the origin of the solvent isotope effect, SIE. However, the

factors mentioned in the theories discussed in this section probably all contribute to the differences in the rates of hydrolysis in light and heavy water. Both the zero point energy effects Bunton and Shiner (61) and the initial and transition state effects of Robertson and Laughton (57) contribute to the SIE. However, at this time comparison of the hydrolysis of compounds in light and heavy water could only serve to provide mechanistic evidence with support from other derived parameters, such as enthalpies, entropies, and heat capacities of activation and from information derived from tracer studies.

HEAT CAPACITIES OF ACTIVATION

The importance of heat capacities of activation was first seriously considered by La Mer (65). He has treated this problem theoretically using a simple unimolecular process for a model. The model considers a system of molecules of index i , which are distributed in various quantum levels having the energies ϵ_i and degeneracies p_i . The molecules in the higher quantum states are indicated by the index j and are assigned undetermined specific individual rate constants. Then the number of molecules, n_j , in the j^{th} quantum state is given by:

$$n_j = \frac{n p_j e^{-\epsilon_j/kT}}{\sum_i p_i e^{-\epsilon_i/kT}}, \quad (9)$$

where k and T have their usual significance.

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The overall rate constant is then

$$k_1 = \frac{-1}{n} \frac{dn}{dt} = \frac{\sum_j k_j p_j e^{-\epsilon_j/kT}}{\sum_i p_i e^{-\epsilon_i/kT}} \quad (10)$$

By taking the natural logarithm of the expression and differentiating, the following expression is obtained:

$$\frac{d \ln k_1}{d T} = \frac{1}{kT^2} \left[\frac{\sum_j \epsilon_j k_j p_j e^{-\epsilon_j/kT}}{\sum_j k_j p_j e^{-\epsilon_j/kT}} - \frac{\sum_i \epsilon_i p_i e^{-\epsilon_i/kT}}{\sum_i p_i e^{-\epsilon_i/kT}} \right] \quad (11)$$

$$\frac{1}{kT^2} \left[\bar{\epsilon} - \bar{\epsilon} \right] = E_{act}/RT^2, \quad (12)$$

where $\bar{\epsilon}$ is the average energy of all the molecules and $\bar{\epsilon}$ is the average energy of the molecules that react. La Mer (65) defined the quantity $N(\bar{\epsilon} - \bar{\epsilon})$ as the energy of activation per mole. This is Tolman's expression (66). The variation of the energy of activation is derived from this expression by further differentiation and the quantity obtained is the heat capacity of activation, C_{act} .

$$kT^2 \frac{d(E_{act})}{dT} = \left(\frac{\sum_j \epsilon_j^2 k_j p_j e^{-\epsilon_j/kT}}{\sum_j k_j p_j e^{-\epsilon_j/kT}} \right) - \left(\frac{\sum_j \epsilon_j k_j p_j e^{-\epsilon_j/kT}}{\sum_j k_j p_j e^{-\epsilon_j/kT}} \right)^2$$

$$- \left(\frac{\sum_i \epsilon_i^2 p_i e^{-\epsilon_i/kT}}{\sum_i p_i e^{-\epsilon_i/kT}} \right) + \left(\frac{\sum_i \epsilon_i p_i e^{-\epsilon_i/kT}}{\sum_i p_i e^{-\epsilon_i/kT}} \right)^2$$

$$\text{or } \frac{d(E_{act})}{dT} = 1/kT^2 \left\{ \left[\overline{\epsilon^2} - (\bar{\epsilon})^2 \right] - \left[\bar{\epsilon}^2 - (\bar{\epsilon})^2 \right] \right\}$$

$$= (\bar{\epsilon}_{j(act)} - \bar{\epsilon}_i) = C_{act} \quad (13)$$

When the heat capacity of activation is neglected, i.e. when equation (13) is equated to zero, the free energy and total energy are indistinguishable. The correction introduced by considering heat capacities can be seen in the thermodynamic expression

$$\int \frac{E_{act} dT}{T^2} = \frac{E_{act}}{RT} + \int \frac{dE_{act}}{dT} d \ln T, \quad (14)$$

where the last term is by definition ΔC_p^\ddagger .

This is the definition of heat capacity which was used by Moelwyn-Hughes in his early determinations of the variation of the energy of activation with temperature. Moelwyn-Hughes stated that the most significant single constant for reactions with a measurable rate is the energy of activation calculated by the method Arrhenius. This, in the majority of cases, is a true constant over the experimental range. He found, however, in the case of the catalysed hydrolysis of certain complicated molecules, the glycosides, that the energy of activation falls markedly with temperature (67) (68) ($dE_A/dT = -98 \pm 4$). This variation in activation energy was found to apply not only to the hydrolysis of the glycosides but also to the hydrolysis of simpler molecules such as the methyl halides. The values obtained by Moelwyn-Hughes (69) for methyl bromide, chloride, and iodide were negative and in the region $-67 \text{ cal. mole}^{-1} \text{ deg.}^{-1}$. He also recognized (69) that to determine the variation of the Arrhenius E_A with temperature requires data of higher accuracy than that generally available in the late 1930's.

This derived C_{act} is also the heat capacity of activation denoted by Moelwyn-Hughes, Robertson, and Sugamori (70) as $\Delta C_p^\ddagger = \left(\frac{\partial E_A}{\partial T} \right)_p$.

However, they have also defined another heat capacity of activation, ΔC_p^\ddagger which is adopted from the thermodynamic expressions of the absolute rate theory (71). The temperature dependence of the rate is expressed in terms of the free energy of activation (ΔG^\ddagger) and Hertzfeld's constant (kT/h)

$$k_1 = (kT/h) \exp (-\Delta G^\ddagger/RT), \quad (15)$$

where k and h are Boltzmann's and Planck's constants, respectively, and ΔG^\ddagger is defined as a free energy of activation. Then the enthalpy of activation can be related to the Arrhenius activation energy by the expression

$$\Delta H^\ddagger = E_A - RT \quad (16)$$

Differentiating eq. 16 with respect to temperature gives:

$$\frac{\partial \Delta H^\ddagger}{\partial T} = \Delta C_p^\ddagger = \Delta C_p^* - R \quad (17)$$

Therefore, values of heat capacities may differ by a constant R .

Moelwyn-Hughes (72) has interpreted the difference in ΔC_p^* values for the hydrolysis tertiary, and secondary and primary halides on the basis of several of his earlier papers in 1938 (69) (72) and has derived a mathematical expression from hypotheses concerning the interaction

energies between one solvent molecule and the solute molecule and between neighbouring solvent molecules. He assumed that the preexponential term of the Arrhenius equation is a function of temperature and, for example, the relaxation frequency which varies inversely as the viscosity of the medium. The apparent energy of activation then contains a term

$$B = -RT^2 \left(\partial \ln \eta / \partial T \right) \quad (18)$$

and ΔC_p^* contains the term dB/dT which is negative and becomes less negative with temperature. Moelwyn-Hughes states that it is the $\partial B/\partial T$ term that creates the difference in ΔC_p^* between the primary and secondary, and the tertiary halides.

Robertson has extensively reviewed (73) solvolysis in water and devotes a section to the interpretation of heat capacities of activation. He states that it was the large changes in ΔH with temperature which could not be accounted for by changes in the substrate during the activation process but must arise from changes in the solvent envelope. It might be assumed from this that the entropy of activation would serve to provide more easily the same information as the heat capacity of activation. Robertson (73) has given several reasons why the ΔC_p^\ddagger would be favoured over ΔS^\ddagger for information concerning the activation process: (a) A differential term has been shown to be a more sensitive

indicator of change and from this point of view ΔC_p^\ddagger would be favoured; (b) ΔS^\ddagger includes steric factors which would have no influence upon ΔC_p^\ddagger ; and (c) the information derived from ΔC_p^\ddagger data already could not have been obtained from the ΔS^\ddagger values obtained for the same reaction. For these reasons Robertson has extensively measured and used the ΔC_p^\ddagger parameter to obtain information concerning the mechanisms of reactions. The point (b) is probably incorrect since the very nature of ΔC_p^\ddagger involves solute-solvent interactions and hence must include some steric factors.

For most reactions the heat capacity of activation is a negative quantity, i.e. as the temperature is increased the energy of activation decreases. Robertson (73) has discussed the hydrolysis of halides primarily in terms of solvation shells. The water molecules which are part of the relatively stable solvent shell about the halide are in equilibrium with those in the bulk solvent. The relative structural stability of water decreases as the temperature increases and a corresponding decrease occurs in the work required to break down the solvent shell which surrounds the solute. Robertson considers that the ΔC_p^\ddagger for the solution process reflects the changes in water - water interaction but the activation process in hydrolysis probably involves only part and that part to an unknown degree.

Robertson and co-workers (73) found that values of ΔC_p^\ddagger for the halides are about -50 cal./mol. deg. and the average value for the sulphonates is about -35 cal./mol. deg. They suggested that this difference of 10-20 cal./mol. deg. was due to interaction of the sulphonic oxygens with the adjacent water molecules which increases the vibrational freedom in the initial state and consequently the free energy required to disrupt the solvent. With the availability of a great deal more data, the possibility of strong ion-water interaction at the transition state and a corresponding contribution to ΔC_p^\ddagger was assumed. Later this interaction was considered to probably be unimportant.

Robertson also discusses the S_N1 and S_N2 mechanisms for hydrolysis in terms of the heat capacity of activation. Experimentally the primary and secondary halides which are believed to hydrolyse by an S_N2 mechanism, have heat capacities of activation in the region of -60 cal./mol. deg. and the tertiary halides reacting by an S_N1 mechanism have ΔC_p^\ddagger values which are considerably more negative. The overall effect in the S_N2 mechanism appears to be determined by the temperature dependence of the work required to break down the solvent shell. The more negative values of ΔC_p^\ddagger for the S_N1 mechanism are also attributed to disruption of the solvent shell but disruption to a greater extent than in the S_N2 case. The explanation for the ΔC_p^\ddagger values measured

is not as simple in all cases and many apparent exceptions to this rule have been obtained and in turn have been explained in terms of destabilization by certain groups (e.g. MeO) in the initial state or by anchimeric assistance of a group during the reaction, e.g. OH. It may be tentatively concluded from this work and the data available that one may distinguish between S_N1 and S_N2 mechanisms by means of ΔC_P^\ddagger values.

The major determinant of ΔC_P^\ddagger , according to Robertson (73), is a water-water interaction in the initial and transition states. Extensive work has been carried out by Kohnstam (74) in mixed solvents containing a large proportion of the organic component and he reaches the conclusion that the factor determining ΔC_P^\ddagger in such cases may be ion-solvent interaction in the transition state.

In the case of S_N1 reactions this difference between ΔC_P^\ddagger values in water and in aqueous organic solvents is proposed to be due to increased solvent-solute interactions by the organic component of the solvent which in turn causes a reduction in the solvent-solvent interactions which are caused by the presence of the solute.

Kohnstam (74) treats the transition state for an S_N2 reaction as quasi-ionic and assumes that the covalent

participation of the water molecule has little effect on ΔC_p^\ddagger and ΔS^\ddagger in aqueous organic solvents. The large values of ΔS^\ddagger observed for hydrolysis of neopentyl bromide and tertiary butyl chloride imply that nearly all the solvent-solute interactions in the initial state disappear in the activated complex. This may arise from the fact that reaction is now occurring by an S_N1 mechanism and the other halides studies are reacting by an S_N2 mechanism which involves an apparent loss of entropy due to the covalent participation of water in the activation process. This idea is consistent with results in aqueous organic solvents but Kohnstam (74) suggests that the interactions proposed by Robertson are probably the controlling factors for the magnitude of the ΔS^\ddagger and ΔC_p^\ddagger parameters in solvolysis reactions in pure water.

Although ΔC_p^\ddagger may be written

$$\Delta C_p^\ddagger = (C^\ddagger - C_p), \quad (19)$$

i.e. it is the difference in the heat capacity of the average substrate molecule between the initial and the transition state, this expression does not provide direct information about the mechanism of the reaction which is being studied. Acceptable physical explanation to account for the experimental results in terms of ΔC_p^\ddagger has not yet been given and again this quantity alone provides little information concerning the mechanism of the reaction being studied.

EXPERIMENTAL

(a) PREPARATION OF MATERIALS

(I) General

Nuclear magnetic resonance spectra were recorded on a Varian A-60 spectrometer and peaks were recorded on the τ scale relative to tetramethylsilane as an internal reference, except in purity determinations of heavy water when acetonitrile was used as an internal standard.

Infrared spectra were recorded on a Perkin-Elmer Model 225 spectrometer using the liquid cell technique with carbon tetrachloride as solvent.

Refractive indices of the ethyl, propyl and tertiary butyl trifluoroacetates were measured on an Abbé refractometer at 20°C using a sodium lamp light source. A Pulfrich refractometer was used to measure the refractive index (n_D^{20}) of methyl trifluoroacetate.

Saponification equivalents were determined by allowing a weighed amount of each ester to react over a period of several days with ca. 250 ml. of conductance water which had been boiled before use to remove dissolved carbon dioxide. The resulting solution was washed into a 1000 ml. volumetric flask which was then filled to the mark with conductance water. This procedure produced a ca. 0.01 N.

solution of the corresponding acid which was titrated against an 0.00986 N. solution of NaOH. The NaOH solution had previously been standardized using weighed samples of dry potassium hydrogen phthalate (BDH, AR grade, not less than 99.9% pure).

Methyl and ethyl trifluoroacetates (Fluka AG purum grade) and acetonitrile (Fisher certified reagent) were distilled using a 12 cm. Vigreux fractionating column.

METHOD 1

Iso-propyl and tertiary butyl trifluoroacetates were prepared by two methods, both of which produced low yields.

Iso-propyl bromide (5.0 g., 0.04 moles) was added to silver trifluoroacetate (0.1 mole excess) dissolved in 125 ml. benzonitrile. The mixture was heated on a steam bath for 30 minutes and the precipitate of silver bromide was then filtered off. The fraction boiling below 100°C was collected and redistilled, using a Vigreux column.

Several attempts to prepare tertiary butyl trifluoroacetate by the same method failed. A second method (75) was modified slightly and produced better yields of both the tertiary butyl and iso-propyl esters.

METHOD 2

Tertiary butyl alcohol (5 g., .068 moles) was added to trifluoroacetic anhydride (18.6 g., 0.3 mole excess) and silver trifluoroacetate (.2 g.). The mixture was re-fluxed gently for 15 minutes after the initial vigorous reaction had subsided. Ice-water (2 ml) was added to the mixture and the aqueous phase was neutralized, using cold 2 N. NaOH, to pH 7-8. The organic phase was separated, dried with $MgSO_4$, filtered and distilled. This method produced tertiary butyl trifluoroacetate in 28% yield and iso-propyl trifluoroacetate in 45% yield.

The physical properties of the four esters are recorded in Table II.

(II) Preparation of Solutions

Light water: Conductance water was obtained by passing tap water through a bed of AR grade Amberlite MB-1 into a Barnsted still where it was distilled, collected and stored in a large sealed flask which had previously been flushed with steam. The still is equipped with a Borosilicate Glass Condensing System and is designed to produce a distillate containing 0.001 ppm. total solid and giving electrical resistances ranging from 1.5 to 5 million ohms per c.c. Water for use in the present work was passed through another deionizing bed of amberlite and collected

TABLE II

PHYSICAL AND CHEMICAL PROPERTIES OF SOME ESTERS DERIVED FROM TRIFLUOROACETIC ACID

Ester	b.p., °C* 760 mm.	% Purity [†]	N.M.R. τ	Refractive Index	$\nu(\text{C}=\text{O})$ cm. ⁻¹
-Me	43 - 43.6 (43)	99.4	6.02	1.29162 (1.29073)	1788 (1786)
-Et	63 - 63.7 (61.3)	98.7	8.58, 5.54	1.3075 (1.30725)	1784
-iPr	71.5 - 72 (73)	97.8	8.64, 4.83	1.3190 (1.3178)	1787
-tBu	83 - 84 (45 @ 145 mm.)	98.9	8.45	1.3345	1782 (1777)

Values in parentheses are literature values.

[†] Calculated from saponification values

* Except where otherwise indicated

The excellent linearity of the Guggenheim plots in cases where supersaturation was avoided indicated that each substrate was kinetically homogeneous.

in a flask which had been rinsed several times with conductance water. Backing electrolyte was prepared by allowing an appropriate amount of ester to hydrolyse in this water reducing the resistance of the water to 10,000 - 40,000 ohms in the cells used as reaction vessels. The cell was filled with backing electrolyte and was placed in the constant temperature bath where it was left, for at least thirty minutes, until the solution reached thermal equilibrium. The ester (enough to produce a 10^{-5} - 10^{-4} M. solution, ca. .6 μ l.) was added directly to the solution in the cell which was stirred and allowed 5 - 15 minutes to become homogeneous before the resistance was measured. In cases where supersaturation posed a problem, a flask containing some backing electrolyte was placed in the bath along with the cell and an appropriate quantity of ester (3 μ l.) was added to the solution in the flask, which was then shaken vigorously for one minute. A 1 ml. aliquot was withdrawn and injected into the cell by means of a syringe. The procedure was then continued as described above.

Heavy Water: Heavy water (99.75% pure) was obtained from Atomic Energy of Canada and solutions were prepared as for light water. After use, the D_2O was collected and redistilled (see Fig. 2) from a small amount of dry AR sodium carbonate. Analysis by N.M.R., using acetonitrile as an

internal reference, indicated that there was no appreciable decrease in purity, and on the completion of this work, the hydrogen content of the heavy water was less than 0.5%.

(b) DESCRIPTION OF APPARATUS

Tinsley Bridge: Figure 3 shows a simplified block diagram of the 4896 Tinsley Bridge with auxiliary equipment. The bridge and accessories were assembled following the description of Robertson (76) with minor modifications. The unknown resistance in the figure refers to the solution contained in the cell. The balance point of the Tinsley Bridge was detected by observing a Hewlett-Packard (H.P.) model 120B oscilloscope.

Oscillator: Power was supplied by a 120 volt line to a H.P. oscillator model 200AB which provided a source of 1000 c/s alternating current to the Tinsley Bridge (Fig. 4). A voltage stabilized power control was not used.

High Gain 1000 c/s Amplifier: The gain of the oscilloscope could be varied to complement the speed of the resistance change. For this purpose a 1000 c/s high gain amplifier was used and was connected to the bridge through the 578-A transformer. The output of the amplifier was fed through shielded leads to the vertical input of the oscilloscope.

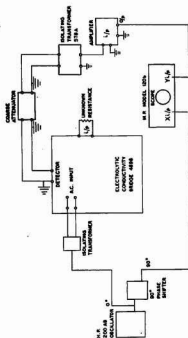


Figure 3
BLOCK DIAGRAM OF CONDUCTIVITY BRIDGE
WAGNER-EARTH SYSTEM

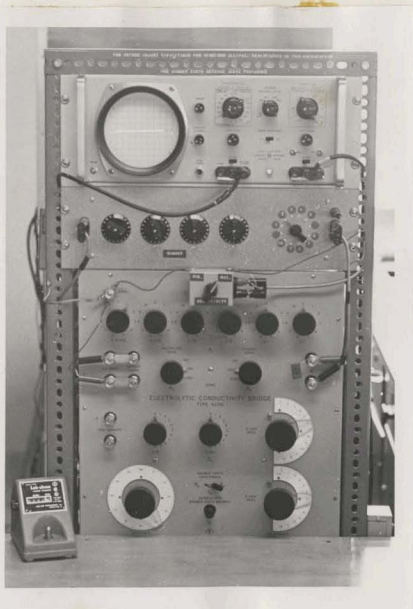


Figure 4

THE TINSLEY A.C. WHEATSTONE BRIDGE

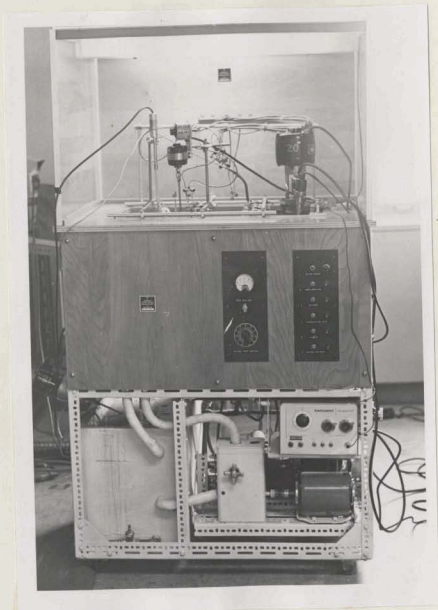
Constant Temperature Bath: The constant temperature bath (Figs. 5 and 6) consisted of a 40" x 29" x 22" metal framed unit supporting a stainless steel tank of dimensions 28" x 13" x 18" which was surrounded by a one inch covering of styrofoam on the outer surface. At all temperatures used in the present study it was necessary to simultaneously heat and cool the water bath to maintain a stable temperature with a deviation of $\pm 0.005^{\circ}\text{C.}$ over a period of several hours. Below the main water bath was another insulated stainless steel tank containing ethylene glycol which was cooled by coils from a Temcuseh $\frac{1}{4}$ h.p. refrigerator. The cooling coils were made of ca. 25 ft. of copper tubing wound in 5 spirals placed ca. 2" apart. By means of a circulating pump the glycol, cooled in this manner, was pumped through copper coils which passed through the water bath. The flow of glycol could be regulated by a control valve in the circulating system. A Fenwal Electronics thermistor (LB21J1) in a Wheatstone bridge circuit was placed in the glycol bath and this provided a rough value for the cooling bath temperature which could be observed on a temperature monitor on the front of the apparatus.

The water in the upper bath was agitated by a Cenco Centrifugal Electric stirrer (circulating capacity, 105 gal. of water per minute) mounted in one corner of the bath. The

bath temperature was regulated by a Sargent "Thermonitor" (S-82050) which had been slightly modified for better performance by changing the voltage across the thermistor-Wheatstone Bridge from 6 to 4 volts. The "Thermonitor" unit consists of a background heater (125 watts) which is controlled by a Variac on the control box, and an intermittent knife heater (250 watts) which is controlled by the thermistor-Wheatstone Bridge circuit coupled to a variable reactor. Photographs of the constant temperature bath are shown in Figs. 5 and 6.

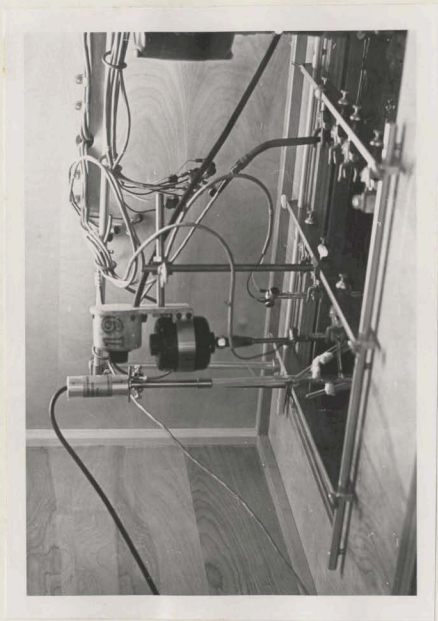
Thermometry: The temperature of the upper thermostated bath was measured by means of a Tinsley Platinum Resistance Thermometer (calibrated by the National Physical Laboratory, Teddington, England) in conjunction with a Mueller Bridge, type 4772, also manufactured by the Tinsley Company. The resistance of the thermometer at zero degrees centigrade (ice point) was checked during the course of the present study using ice obtained from several sources: fractionally frozen conductance water, crushed ice cubes prepared from conductance water and crushed ice obtained directly from a "Scotsman" ice producing machine. In each case the ice was well mixed with conductance water so that it did not float in a 1 litre Dewar flask. The thermometer was placed carefully in the ice, was allowed at least 15 minutes for thermal equilibration and

Figure 5



THERMOSTAT

Figure 6

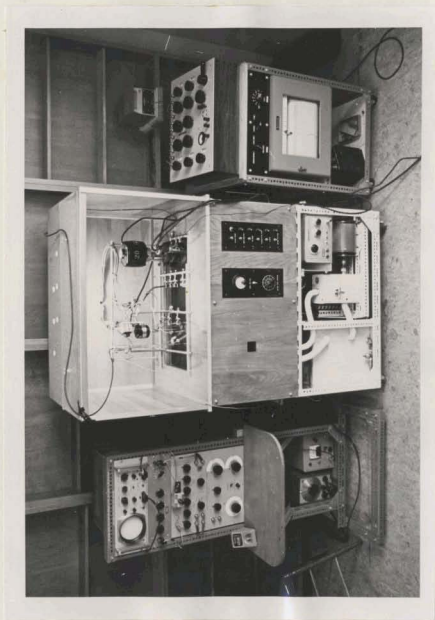


CLOSE-UP OF THERMOSTAT BATH

resistance readings were obtained until there was no apparent change in temperature. Using this method the fractionally frozen conductance water and the crushed ice from the machine produced identical results, and it was found that the ice point, R_0 , had changed from 24.923 ohms to 24.925 ohms. This comparison of ice from different sources indicates that ice from the "Scotsman" machine, which was most easily procured, was sufficiently pure to determine a new ice point for the Platinum Resistance Thermometer. The slightly higher value obtained using crushed ice cubes can be attributed to impurities introduced during handling which was necessarily greater than in the other cases. The new R_0 from the ice point calibration was used in subsequent temperature determinations.

A Honeywell recorder with an attached thermistor was used to record changes in temperature which occurred during a run and also proved useful in indicating when the temperature of the bath had become stable after a change from one temperature to another. A Fenwal thermistor served as one arm in a Wheatstone Bridge circuit and five resistors were used to control the sensitivity of the recorder. Ten small divisions on the recorder chart paper were found to be equivalent to 0.005°C . at maximum sensitivity. Fig. 7 shows the Tinsley and Mueller bridges, the Honeywell recorder and the constant temperature bath.

Figure 7



COMPLETE APPARATUS USED FOR RATE DETERMINATIONS

Cells: Three stirred cells similar to those previously described by Hyne and Robertson (77) were used as reaction vessels. A Photograph of one of the cells appears in Figure 8. The cells were filled with light water backing electrolyte and several experiments were carried out in them until the rates measured agreed to $\pm 0.3\%$ at the same temperature. The reaction vessels were then "conditioned" and could be used for the required experiments in light water. Before the heavy water experiments were carried out the cells were dried in an oven at 100°C and the conditioning procedure was carried out as described above using heavy water. Stirring of the solution in the cells was effected by means of a bar magnet which was attached to the end of a stirrer shaft immersed in the bath. The cell was stirred during the entire run and the speed of rotation was controlled by a variable resistor on the stirrer motor.

The approximate cell constants were determined by comparing the resistance of a solution in a cell of known cell constant with the resistance of the same solution in the cell of unknown cell constant. The conductance of a solution may be represented by:

$$\Lambda_t = \frac{\kappa \cdot 10^{-3}}{R_t \cdot C'_t} \quad (20)$$

where κ is the cell constant of the vessel, R , the resistance,

Figure 8

CONDUCTANCE CELLS



and C'_t the concentration of the solution at a given temperature. For a given solution C'_t and Λ_t are constants independent of the cell used. Therefore, for cells, 1, 2, 3, and 4, the equation may be written

$$\kappa_1/R_{t1} = \kappa_2/R_{t2} = \kappa_3/R_{t3} = \kappa_4/R_{t4} \quad (21)$$

If κ_4 is known, then κ_1 , κ_2 and κ_3 may be determined.

(c) GUGGENHEIM-ROBERTSON METHOD

The method proposed by Guggenheim (78) has proved useful for solvolytic rate determinations in instances when the acid related to the leaving group is strong. However, the validity of this method may be questioned in the case of weaker acids such as trifluoroacetic acid which has a pK_a of 0.23 (21). In this work it was decided to investigate quantitatively the effect that the decrease in acid strength would have upon rate constants determined by the Guggenheim-Robertson Method (GRM). The following is a derivation of the expression used for this purpose.

For a pseudo first order reaction, e.g. the solvolysis of a halide, the rate expression may be written in the following manner:

$$C_t = C_0 e^{-kt} \quad (22)$$

where C_0 is the initial concentration of substrate and C_t is

the concentration of substrate at time t . In order to know C_t the concentration of C_0 must be known and this cannot always be obtained either easily or accurately. The first order rate constant can be obtained from a plot of $\log C_t$ vs. t if C_0 is known.

The GRM eliminates the C_0 term and, consequently, unnecessary work. The equation proposed by Guggenheim (78) for first order rate determinations is of the form:

$$C_{t+\tau} - C_t = C_0 e^{-kt} (1 - e^{-k\tau}) \quad (23)$$

where C_0 is the initial concentration of the substrate, k is the rate constant, C_t is the concentration of the product, an acid in this case, at time t , and $C_{t+\tau}$ is the concentration at time $t + \tau$; τ is the time interval greater than two times the half-life of the reaction. A set of concentration and time readings are taken for a half-life of the reaction and the reaction is then left for a half-life. Measurements are resumed and a corresponding concentration reading is taken τ minutes after the first resistance was measured (see Fig. 9).

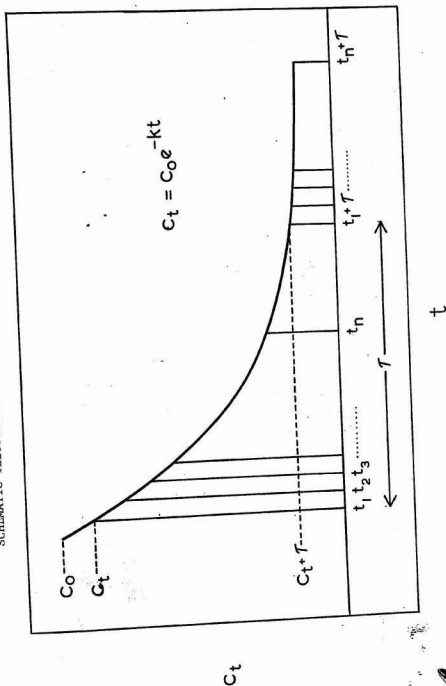
Ives (79) has proposed that for a dilute solution of a weak electrolyte the equivalent conductance may be represented in the form:

$$1/\Lambda \approx 1/\Lambda_0 + AC/K_c \Lambda_0^2 \quad (24)$$

which neglects the effect of ionic strength on both the

Figure 9

SCHEMATIC ILLUSTRATION OF THE GUGGENHEIM METHOD



activity and mobility of the ions. The quantity is the equivalent conductance of the weak electrolyte of concentration C , Λ_0 is the limiting equivalent conductance and K_c is the classical dissociation constant which has not been corrected for the variation of activities or mobilities with the ionic strength. The equivalent conductance in terms of the experimental parameters is given by the equation:

$$\Lambda = \kappa 10^3 / RC \quad (25)$$

where κ is the cell constant of the reaction vessel and R is the resistance of the solution.

Substituting for Λ in Equation (24) gives

$$RC / \kappa 10^3 = 1 / \Lambda_0 + \kappa 10^3 / K_c \Lambda_0^2 R \quad (26)$$

and after some manipulation

$$C = \frac{\kappa 10^3}{\Lambda_0} \cdot \frac{1}{R} - \frac{\kappa^2 10^6}{\Lambda_0^2 K_c R^2} \quad (27)$$

or

$$C_t = P / R_t + P^2 / K_c R_t^2 \quad (28)$$

where $P = \kappa 10^3 / \Lambda_0$ which is a constant for any particular reacting system. The calculation of P for trifluoroacetic acid has been carried out in this study. Λ and K_c were estimated by an extrapolation of data reported by Henne and Fox (80). Assuming that $\log K$ vs. $1/T$ and Λ_0 vs. T were

linear plots, values of K and Λ_0 at 15°C were obtained. The cell constant was determined in a manner described on page 59. Values for these constants are as follows:

$$\Lambda_{15} = 332, \quad K_{15} = .652, \quad \text{and } \kappa = .196604$$

The basic assumption of the GRM is that the concentration of the strong acid is proportional to the reciprocal of the resistance, i.e. that the squared term in Equation (28) is negligible. Robertson (73) has compared rate constants calculated using this assumption to rate constants obtained by two other methods for the same reaction. He found empirically that the differences in rates derived by the Guggenheim method, combined with the Onsager equation and the GRM, as outlined above, were minor and the differences in the derived thermodynamic parameters were also unimportant.

Substituting Equation (28) into the Guggenheim expression (Equation 23) produces the following equation:

$$P/R_{t+\tau} + P^2/KR_{t+\tau}^2 = (P/R_t + P^2/KR_t^2) = C_0 e^{-kt} (1 - e^{-k\tau})$$

$$P(1/R_{t+\tau} - 1/R_t) + P^2/D_C (1/R_{t+\tau}^2 - 1/R_t^2) = C_0 e^{-kt} (1 - e^{-k\tau})$$

$$(1/R_{t+\tau} - 1/R_t) (1/P/K_C (1/R_{t+\tau} - 1/R_t)) = \frac{-C_0 (1 - e^{-k\tau}) e^{-kt}}{P} \quad (29)$$

For any given experiment the factor $C_0(1 - e^{-k\tau})/P$ is a constant, Q , say. Therefore,

$$(1/R_{t+\tau} - 1/R_t) (1 + P/K_C (1/R_{t+\tau} + 1/R_t)) = Q e^{-kt} \quad (30)$$

By taking logs of Equation (30) the following expression is derived:

$$\log (1/R_{t+\tau} - 1/R_t) + \log (1 + P/K_C (1/R_{t+\tau} + 1/R_t)) = \log Q - kt/2.3026 \quad (31)$$

Clearly from Equation (31) a plot of $\log (1/R_{t+\tau} - 1/R_t)$ vs. time will only be a straight line if the second term on the right hand side of the equation is small, i.e. if $P/K_C (1/R_{t+\tau} + 1/R_t)$ approaches zero.

To demonstrate that this term is, in fact, negligible, a run was chosen at random for further study. As shown by column five in Table III the term being considered is very small and essentially constant, and its inclusion in the calculations was found to make no difference to the rate constant.

(d) AUTOCATALYSIS

Since the solvolysis of trifluoroacetates is known to be acid catalysed (1), the possibility that autocatalysis might be occurring in this case was examined. The GRM requires that the reaction be followed for at least three half-lives to determine a rate constant since R_t and $R_{t+\tau}$ are

TABLE III

A TYPICAL RUN - HYDROLYSIS OF ETHYL TRIFLUOROACETATE AT 14.95°C

Time (sec)	R_t	$R_{t+\tau}$	$1/R_t - 1/R_t$ = $D(\times 10^5)$	$1 + P/K(1/R_t + 1/R_t)$ = M	log D	log D + log M
0	7710.0	4173.0	10.993	1.0003	4.0401	4.0402
41	7170.0	4137.5	10.222	1.0003	4.0095	4.0096
80	6760.0	4107.0	9.5558	1.0004	3.9803	3.9805
119	6410.0	4079.5	8.9122	1.0004	3.9500	3.9502
161	6100.0	4052.2	8.2845	1.0004	3.9183	3.9185
202	5840.0	4027.4	7.7066	1.0004	3.8869	3.8871
240	5630.0	4006.2	7.1993	1.0004	3.8573	3.8575
280	5440.0	3985.8	6.7067	1.0004	3.8265	3.8267
321	5270.0	3966.6	6.2352	1.0004	3.7949	3.7951
360	5130.0	3949.4	5.8271	1.0004	3.7655	3.7657
401	5000.0	3933.0	5.4259	1.0004	3.7345	3.7347

followed over a half-life each and ideally there should be a period of one half-life between them, so that $\tau \approx 2t_{1/2}$. This means that at least 87.5% of the reaction must be followed in any given run. However, in this study, the time of one of the experiments was extended to include four half-lives. Four sets of resistance readings, separated by a brief interval, were taken at identical times. The results and calculations are given in Table IV and the Guggenheim plots are shown in Figure 10. The three rate constants calculated are (a) 17.62×10^{-4} , (b) 17.58×10^{-4} , and (c) 17.66×10^{-4} and the average rate constant obtained from the three possible determinations is $17.62 \pm 0.04 \times 10^{-4} \text{ s}^{-1}$. The deviation is well within the experimental reproducibility, indicating that for 93.3% of the reaction (four half-lives) there is no important acid-catalysis by the hydronium ion formed in the reaction.

(e) RATE CONSTANTS

The rate constants which have been measured for the solvolysis of the methyl, ethyl, iso-propyl and tertiary butyl trifluoroacetates in water and deuterium oxide are recorded in Tables V, VI, and VII along with the calculated rates and deviations. At each temperature a minimum of three (usually four) rate determinations were carried out and the root mean square deviation between the rates was less than $\pm 0.4\%$. Due

Figure 10

FIRST ORDER PLOTS FOR THE RATES
OF HYDROLYSIS OF CF_3COOEt

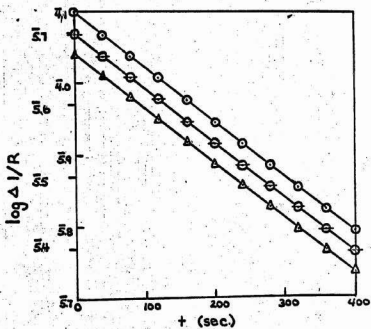


TABLE IV

DATA FOR EXTENDED GUGGENHEIM PLOT (OVER FOUR HALF-LIVES)

HYDROLYSIS OF ETHYL TRIFLUOROACETATE AT 14.95°C

Time (sec)	R_1	R_2	R_3	R_4	$\Delta_{13}/1R \times 10^5$
0	7710.0	4873.0	4173.0	3920.0	10.993
41	7170.0	4770.0	4137.5	3905.6	10.222
80	6760.0	4683.0	4107.0	3893.0	9.5558
119	6410.0	4604.0	4079.5	3881.3	8.9122
161	6100.0	4527.0	4052.2	3870.0	8.2845
202	5840.0	4461.0	4027.4	3859.1	7.7066
240	5630.0	4404.0	4006.2	3850.2	7.1993
280	5440.0	4351.0	3985.8	3841.4	6.7067
321	5270.0	4301.0	3966.6	3832.8	6.2352
360	5130.0	4257.0	3949.4	3825.5	5.8271
401	5000.0	4215.0	3933.1	3818.2	5.4259

(Continued)

TABLE IV (Continued)

Time (sec)	$\Delta_{24}/1R \times 10^5$	$\Delta_{14}^1/R \times 10^5$	$\log \Delta_{13}^1/R$	$\log \Delta_{24}^1/R$	$\log \Delta_{14}^1/R$
0	4.9890	12.540	4.0401	5.6980	4.0983
41	4.6399	11.657	4.0095	5.6665	4.0666
80	4.333	10.894	5.9803	5.6368	4.0372
119	4.0443	10.164	5.9500	5.6068	4.0071
161	3.7501	9.4464	5.9183	5.5740	5.9753
202	3.4963	8.7895	5.8869	5.5436	5.9440
240	3.2660	8.2107	5.8573	5.5140	5.9144
280	3.0490	7.6498	5.8265	5.4842	5.8837
321	2.8402	7.1153	5.7949	5.4533	5.8522
360	2.6497	6.6472	5.7655	5.4232	5.8226
401	2.4656	6.1904	5.7345	5.3919	5.7917

to the low solubility of the esters in water, supersaturation occurred in several instances. The effect of supersaturation was to produce an apparently lower rate constant or a curved instead of a linear plot. Such runs were not included in the calculations. An improved method, (described earlier) of introducing the ester into the cell, eliminated these difficulties.

The temperature dependence of the rates in the case of the ethyl ester could be represented within the experimental error by the equation:

$$\log k = A/T + B \log T + C \quad (32)$$

The parameters, A, B, and C were derived by a least squares determination on an IBM 1620 computer. Using the convention adopted by Robertson (73) the thermodynamic parameters ΔH_O^\ddagger , ΔS_O^\ddagger , ΔC_P^\ddagger , ΔH_t^\ddagger and ΔS_t^\ddagger could be calculated from the empirical parameters (A, B, and C) above. These are recorded in Table VIII.

An overall picture of the solvolysis of the four esters was obtained by carrying out a less extensive study of the three other esters in light and heavy water and the ethyl ester in heavy water. The temperature dependence of the methyl, iso-propyl and tertiary butyl esters could be

TABLE V

RATES OF SOLVOLYSIS OF SOME TRIFLUOROACETATES IN WATER

Ester	Temperature	k(exp)	k(calc)	k(exp) - k(calc)
	°C	$\times 10^4$	$\times 10^4$	$\times 10^4$
CF ₃ COOMe	4.979	22.99	23.02	- .03
	4.979	23.15	23.02	+ .13
	4.979	23.01	23.02	- .01
	4.979	23.20	23.02	+ .18
	9.958	32.84	32.83	+ .01
	9.958	32.77	32.83	- .06
	9.958	32.74	32.83	- .09
	9.958	32.86	32.83	+ .03
	14.949	46.33	46.33	.00
	14.949	46.57	46.33	+ .24
	14.949	46.36	46.33	+ .03
CF ₃ COOEt	4.977	8.761	8.790	- .029
	4.977	8.781	8.790	- .099
	4.977	8.754	8.790	- .036
	4.977	8.811	8.790	+ .021
	9.956	12.52	12.51	+ .01
	9.956	12.53	12.51	+ .02
	9.956	12.53	12.51	+ .02
	9.956	12.49	12.51	- .02
	9.956	12.60	12.51	+ .09

(Continued)

TABLE V (Continued)

Ester	Temperature °C	k(exp) $\times 10^4$	k(calc) $\times 10^4$	k(exp) - k(calc) $\times 10^4$
CF ₃ COOEt				
(Continued)	14.949	17.59	17.61	- .02
	14.949	17.49	17.61	- .12
	14.949	17.61	17.61	.00
	14.949	17.58	17.61	- .03
CF ₃ COOiPr	4.979	2.422	2.442	- .020
	4.979	2.435	2.442	- .007
	4.979	2.440	2.442	- .002
	4.979	2.436	2.442	- .006
	9.958	3.481	3.474	+ .007
	9.958	3.496	3.474	+ .022
	9.958	3.505	3.474	+ .031
	14.951	4.888	4.883	+ .005
	14.951	4.845	4.883	- .038
	14.951	4.877	4.883	- .006
	14.951	4.845	4.883	- .038
CF ₃ COOtBu	4.977	.5757	.5716	+ .0041
	4.977	.5709	.5716	- .0007
	4.977	.5703	.5716	- .0013
	9.958	1.292	1.306	- .018
	9.958	1.297	1.306	- .009
	9.958	1.292	1.306	- .018

(Continued)

TABLE V (Continued)

Ester	Temperature °C	k(exp)	k(calc)	k(exp) - k(calc)
		$\times 10^4$	$\times 10^4$	$\times 10^4$
CF ₃ COOtBu	14.951	2.922	2.912	+ .010
(continued)	14.951	2.902	2.912	- .010
	14.951	2.922	2.912	+ .010

TABLE VI
RATES OF SOLVOLYSIS OF SOME TRIFLUOROACETATES
IN DEUTERIUM OXIDE

<u>Ester</u>	<u>Temperature</u> <u>°C</u>	<u>k(exp)</u> <u>x 10⁴</u>	<u>k(calc)</u> <u>x 10⁴</u>	<u>k(exp) - k(calc)</u> <u>x 10⁴</u>
CF ₃ COOMe	4.975	6.509	6.511	- .002
	4.975	6.466	6.511	- .045
	4.975	6.484	6.511	- .027
	4.975	6.486	6.511	- .025
	9.956	9.425	9.358	.067
	9.956	9.398	9.358	.040
	9.956	9.402	9.358	.044
	9.956	9.456	9.358	.098
	14.949	13.23	13.30	- .07
	14.949	13.31	13.30	+ .01
	14.949	13.23	13.30	+ .07
	14.949	13.20	13.30	- .10
CF ₃ COOEt	4.975	2.383	2.398	- .015
	4.975	2.384	2.398	- .014
	4.975	2.391	2.398	- .007
	4.975	2.403	2.398	.005

(Continued)

TABLE VI (Continued)

Ester	Temperature	k(exp)	k(calc)	k(exp) - k(calc)
	°C	$\times 10^4$	$\times 10^4$	$\times 10^4$
	9.956	3.475	3.477	- .002
	9.956	3.478	3.477	.001
	9.956	3.457	3.477	- .020
	9.956	3.462	3.477	- .015
	14.946	4.978	4.977	.001
	14.946	4.941	4.977	- .036
	14.946	4.944	4.977	- .033
	14.946	4.960	4.977	- .017
	4.979	.6356	.6379	- .0023
	4.979	.6351	.6379	- .0028
	4.979	.6315	.6379	- .0064
	4.979	.6334	.6379	- .0045
	9.956	.9176	.9178	- .0002
	9.956	.9210	.9178	.0032
CF ₃ COOiPr	9.956	.9237	.9178	.0059
	9.956	.9172	.9178	- .0006
	9.956	.9176	.9178	- .0002
	14.949	1.298	1.305	.007
	14.949	1.296	1.305	.009
	14.949	1.297	1.305	.008
	14.949	1.293	1.305	.012

(Continued)

TABLE VI (Continued)

Ester	Temperature	k(exp)	k(calc)	k(exp) - k(calc)
	°C	$\times 10^4$	$\times 10^4$	$\times 10^4$
CF ₃ COOtBu	4.977	.4657	.4614	.0043
	4.977	.4599	.4614	- .0015
	4.977	.4613	.4614	- .0001
	9.958	1.085	1.076	.009
	9.958	1.075	1.076	- .001
	9.958	1.075	1.076	- .001
	9.958	1.069	1.076	- .007
	14.951	2.458	2.447	.011
	14.951	2.445	2.447	- .002
	14.951	2.447	2.447	.000

TABLE VII

RATES OF SOLVOLYSIS OF ETHYL TRIFLUOROACETATE IN WATER

Temperature °K	k(exp) $\times 10^4$ (sec ⁻¹)	k(calc) $\times 10^4$ (sec ⁻¹)	k(exp) - k(calc) $\times 10^4$
273.341	6.116	6.093	+ .023
273.341	6.155	6.093	+ .062
273.341	6.090	6.093	- .003
273.341	6.120	6.093	+ .027
274.255	6.542	6.545	- .003
274.255	6.514	6.545	- .031
274.255	6.502	6.545	- .043
274.255	6.539	6.545	- .006
275.544	7.200	7.230	- .030
275.544	7.200	7.230	- .030
275.544	7.188	7.230	- .042
275.544	7.176	7.230	- .054
275.544	7.200	7.230	- .030
276.162	7.568	7.579	- .011
276.162	7.549	7.579	- .030
276.162	7.568	7.579	- .011
276.162	7.576	7.579	- .003

(Continued)

TABLE VII (Continued)

Temperature °K	k(exp) $\times 10^4 \text{ (sec}^{-1}\text{)}$	k(calc) $\times 10^4 \text{ (sec}^{-1}\text{)}$	k(exp) - k(calc) $\times 10^4$
277.032	8.098	8.094	+ .004
277.032	8.151	8.094	+ .057
277.032	8.124	8.094	+ .030
277.032	8.155	8.094	+ .061
277.032	8.140	8.094	+ .046
278.135	8.761	8.788	- .027
278.135	8.761	8.788	- .027
278.135	8.754	8.788	- .034
278.135	8.811	8.788	+ .023
279.141	9.441	9.463	- .022
279.141	9.483	9.463	+ .020
279.141	9.537	9.463	+ .074
279.141	9.494	9.463	+ .031
280.104	10.15	10.15	.00
280.104	10.12	10.15	- .03
280.104	10.15	10.15	.00
280.104	10.14	10.15	- .01
281.135	10.91	10.93	- .02
281.135	10.84	10.93	- .09
281.135	10.95	10.93	+ .02
281.135	10.91	10.93	- .02

(Continued)

TABLE VII (Continued)

Temperature °K	k(exp) $\times 10^4 \text{ (sec}^{-1}\text{)}$	k(calc) $\times 10^4 \text{ (sec}^{-1}\text{)}$	k(exp) - k(calc) $\times 10^4$
282.125	11.80	11.72	+ .08
282.125	11.77	11.72	+ .05
282.125	11.73	11.72	+ .01
282.125	11.72	11.72	.00
283.113	12.52	12.56	- .04
283.113	12.53	12.56	- .03
283.113	12.53	12.56	- .03
283.113	12.49	12.56	- .07
283.113	12.60	12.56	+ .04
284.167	13.52	13.50	+ .02
284.167	13.48	13.50	- .02
284.167	13.42	13.50	- .08
284.167	13.55	13.50	+ .05
285.148	14.51	14.44	+ .07
285.148	14.56	14.44	+ .12
285.148	14.53	14.44	+ .09
285.148	14.55	14.44	+ .11
286.110	15.38	15.40	- .02
286.110	15.32	15.40	- .08
286.110	15.45	15.40	+ .05
286.110	15.41	15.40	+ .01

(Continued)

TABLE VII (Continued)

Temperature °K	k(exp) $\times 10^4$ (sec ⁻¹)	k(calc) $\times 10^4$ (sec ⁻¹)	k(exp) - k(calc) $\times 10^4$
287.141	16.55	16.49	+ .06
287.141	16.51	16.49	+ .02
287.141	16.47	16.49	- .02
287.141	16.57	16.49	+ .08
288.107	17.59	17.57	+ .02
288.107	17.49	17.57	- .08
288.107	17.61	17.57	+ .04
288.107	17.58	17.57	+ .01
289.137	18.65	18.78	- .13
289.137	18.81	18.78	+ .03
289.137	18.69	18.78	- .09
289.137	18.66	18.78	- .12
290.123	20.09	20.00	+ .09
290.123	19.97	20.00	- .03
290.123	20.10	20.00	+ .10
290.123	19.98	20.00	- .02
291.171	21.32	21.36	- .04
291.171	21.34	21.36	- .02
291.171	21.38	21.36	+ .02
291.171	21.20	21.36	- .16

(Continued)

TABLE VII (Continued)

Temperature °K	k(exp) $\times 10^4$ (sec ⁻¹)	k(calc) $\times 10^4$ (sec ⁻¹)	k(exp) - k(calc) $\times 10^4$
292.158	22.78	22.71	+ .07
292.158	22.83	22.71	+ .12
292.158	22.76	22.71	+ .05
292.158	22.89	22.71	+ .18
292.158	22.76	22.71	+ .05
293.179	24.14	24.18	-.04
293.179	24.11	24.18	-.07
293.179	24.24	24.18	+ .06
293.179	24.16	24.18	-.02
294.208	25.51	25.73	-.22
294.208	25.57	25.73	-.16
294.208	25.61	25.73	-.12
294.208	25.51	25.73	-.22
295.184	27.37	27.27	+ .10
295.184	27.18	27.27	-.09
295.184	27.38	27.27	+ .11
295.184	27.47	27.27	+ .20

adequately represented by the following equation:

$$\log (k/T) = A/T + B \quad (33)$$

for reactions in both light and heavy water. Despite the more extensive study mentioned earlier, for comparative purposes the results obtained for the hydrolysis of the ethyl ester at ca. 5, 10 and 15°C have been treated in the same way as those obtained for the other esters. The values of the empirical parameters, A and B, were determined by a least squares method and the thermodynamic parameters ΔH^\ddagger and ΔS^\ddagger were calculated from the Eyring equation which gives

$$A = \Delta H^\ddagger / 2.3026 R \quad (34)$$

$$\text{and } B = \Delta S^\ddagger / 2.3026 R + \log (k/h) \quad (35)$$

The values for these thermodynamic parameters for solvolysis of the esters in both light and heavy water are reported in Table IX. By means of equation (33), rates in light and heavy water at 5, 10, and 15°C were determined and Table X records these rates along with the solvent isotope effects at these temperatures. The results in this table have an accuracy of about $\pm 1.0\%$.

Least squares A, B and C values derived from the data and appropriate to equations (32) and (33) are recorded in Table XI.

TABLE VIII

THERMODYNAMIC PARAMETERS FOR ETHYL TRIFLUOROACETATE IN WATER

Temperature t_K	ΔH_t^\ddagger (Kcal. mole ⁻¹)	ΔS_t^\ddagger (cal. mole ⁻¹ deg. ⁻¹)
273.16	11.16	-32.2
298.16	9.486	-38.1
323.16	7.811	-43.5
348.16	6.137	-48.5
373.16	4.462	-53.1

$$\Delta H_o^\ddagger = 29.459$$

$$\Delta S_o^\ddagger = 343.56$$

$$\Delta C_{po}^\ddagger = -66.987$$

TABLE IX

THERMODYNAMIC PARAMETERS (AT 10°C) FOR THE SOLVOLYSIS
OF SOME TRIFLUOROACETATES IN H₂O AND D₂O

CF ₃ COOR (-R)	ΔH^\ddagger (H ₂ O) (Kcal. mole ⁻¹ .)	ΔH^\ddagger (D ₂ O) (Kcal. mole ⁻¹ .)	ΔS^\ddagger (H ₂ O) (cal. mole ⁻¹ deg. ⁻¹ .)	ΔS^\ddagger (D ₂ O) (cal. mole ⁻¹ deg. ⁻¹ .)
-Me	10.60 ± 0.03	10.84 ± 0.05	-32.3 ± 0.3	-34.0 ± 0.4
-Et	10.53 ± 0.03	10.08 ± 0.03	-34.5 ± 0.3	-35.1 ± 0.3
-iPr	10.51 ± 0.05	10.87 ± 0.05	-37.1 ± 0.6	-38.5 ± 0.7
-tBu	25.44 ± 0.06	26.08 ± 0.05	13.7 ± 0.5	15.5 ± 0.0

TABLE X

SOLVENT ISOTOPE EFFECT FOR THE SOLVOLYSIS OF TRIFLUOROACETATES

IN WATER AND DEUTERIUM OXIDE AT 5, 10 AND 15°C

Ester	T °C	$k(\text{H}_2\text{O}) \times 10^4 \text{ (sec}^{-1}\text{)}$	$k(\text{D}_2\text{O}) \times 10^4 \text{ (sec}^{-1}\text{)}$	$k(\text{H}_2\text{O})/k(\text{D}_2\text{O})$
CF_3COOMe	5.00	23.05	6.522	3.53
CF_3COOMe	10.00	32.93	9.386	3.51
CF_3COOMe	15.00	46.45	13.31	3.49
CF_3COOEt	5.00	8.800	2.403	3.66
CF_3COOEt	10.00	12.54	3.485	3.60
CF_3COOEt	15.00	17.66	4.991	3.54
CF_3COOiPr	5.00	2.445	.6384	3.83
CF_3COOiPr	10.00	3.481	.9198	3.78
CF_3COOiPr	15.00	4.899	1.309	3.74
CF_3COOtBu	5.00	.5748	.4627	1.24
CF_3COOtBu	10.00	1.319	1.084	1.22
CF_3COOtBu	15.00	2.942	2.465	1.19

TABLE XI

EMPIRICAL PARAMETERS FOR THE SOLVOLYSIS OF SOME TRIFLUOROACETATES
IN WATER AND DEUTERIUM OXIDE

<u>Ester</u>	<u>Solvent</u>	<u>A</u>	<u>B</u>	<u>C</u>
CF ₃ COOMe	H ₂ O	-2316.9	3.2476	
CF ₃ COOEt	H ₂ O	-2302.6	2.7782	
CF ₃ COOiPr	H ₂ O	-2297.4	2.2033	
CF ₃ COOtBu	H ₂ O	-5561.5	13.309	
CF ₃ COOMe	D ₂ O	-2368.3	2.8843	
CF ₃ COOEt	D ₂ O	-2422.0	2.6438	
CF ₃ COOiPr	D ₂ O	-2376.0	1.9026	
CF ₃ COOtBu	D ₂ O	-5700.1	13.713	
CF ₃ COOEt	D ₂ O	-6439.0	-32.714	100.0558

DISCUSSION

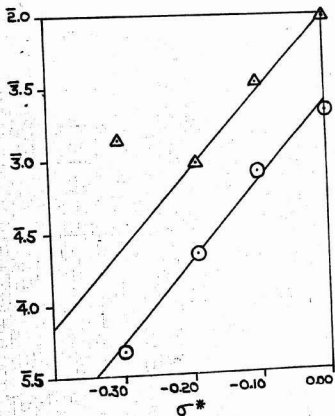
DISCUSSION

The purpose of this discussion will be to attempt an evaluation of the information derived from the rates of reaction, the thermodynamic parameters, and the solvent isotope effects for the hydrolysis of the trifluoroacetates which form the experimental basis for this thesis.

Since the development of the transition-state theory by Eyring (81) and Evans and Polanyi (82), it has been considered that the free energy of activation is the most useful parameter for the discussion of the influence of structure on reactivity (81)(82). The present series of compounds seem to offer an exception to this rule.

It is evident in the case of the trifluoroacetates studied that, depending on the temperature at which the reaction is considered, a treatment based solely on free energy relationships would be ambiguous. Thus, Figure 11 shows that at 3.37°C the order of reactivity of the esters is $\text{Me} > \text{Et} > i\text{Pr} > t\text{Bu}$ and a reasonable Taft correlation is obtained. This leads to the conclusion that the esters all react by the same mechanism and that the reactivity sequence is dominated by the inductive effect of the alkyl group. Figure 11 also shows, however, that at 25°C the Taft correlation breaks down for the $t\text{Bu}$ ester and the reactivity order is changed to $\text{Me} > \text{Et} > t\text{Bu} > i\text{Pr}$.

Correlation of rates of hydrolysis of
alkyl trifluoroacetates with the
Taft σ^* parameter.



Δ Extrapolated rates at 25°C.

\odot Extrapolated rates at 3.73°C.

The order of reactivity at 3.73°C is typical of S_N2 processes but compared to other well established S_N2 reactions the reactivity range is small (83). The reactivity sequence at 25°C is typical of a reaction series where there is a change in mechanism $S_N1 \rightarrow S_N2$ in the region of the secondary ester (16). However, the range of reactivity is again small.

It has been previously mentioned in this thesis that $H_2^{18}O$ studies indicate that the methyl trifluoroacetate is hydrolysed in water by a $B_{AC}2$ mechanism. Figure 1 shows a correlation between the rates of hydrolysis in pure water of a series of ethyl esters derived from a sequence of halogenoacetic acids and the pK_a 's of the related acids. The rates increase as the pK_a 's decrease suggesting that nucleophilic attack at the carbonyl carbon is the main driving force for the ethyl esters' hydrolyses and it may be concluded from this correlation that the ethyl trifluoroacetate also reacts by a $B_{AC}2$ mechanism.

In the present work the change in mechanism referred to earlier is probably from the spontaneous ester hydrolysis mechanism, $B_{AC}2$, to an S_N1 alkyl oxygen fission process. However, this does not mean that in going from the hydrolysis of the secondary to the tertiary ester that there is necessarily an abrupt change from the $B_{AC}2$ type to the S_N1

type reaction. It is conceivable, in fact, that the hydrolysis of both esters takes place by a mixed mechanism and that in the case of the tertiary ester, the reaction is dominated by a carbonium ion intermediate formation. Further, if the idea of a mixed mechanism is accepted, there is a possibility in the case of the tertiary tri-fluoroacetate that the $B_{Ac}2$, S_N1 , and S_N2 mechanisms are all contributing to the reaction rate.

However, from the relative rates and free energies of activation only, it seems impossible to obtain information that is not mechanistically ambiguous. It is therefore necessary to scrutinize other parameters to throw further light on the mechanism.

Enthalpies and Entropies of Activation

In a recent review Schallager and Long (84) have pointed out that the problem of reaction mechanism can be approached from the standpoint of entropies of activation in two different ways. On the one hand the mechanistic significance may be attached to ΔS^\ddagger by comparing the measured value of the parameter with that expected for a proposed transition-state structure by calculation of appropriate partition functions. On the other hand, a change in mechanism of two similar reactants is sometimes indicated by two very different entropies of activation.

In the present work it appears, using the latter approach, that both the entropies and enthalpies of activation fall into two distinct categories: the primary and secondary esters which have entropies of activation which are essentially constant in the region of -35 e.u.'s, and the enthalpies of activation which are relatively constant at 10.5 Kcal/mole. The tertiary ester provides the second category since in its case the enthalpy of activation is 15 Kcal/mole higher than that for the primary and secondary esters and the entropy of activation is ca. 50 e.u.'s more positive.

These observations are consistent with the $B_{Ac}2$ mechanism for the primary and secondary esters and an S_N1 alkyl oxygen fission or a mixed mechanism for the tertiary ester. From the review by Schallager and Long (84) and other discussions of thermodynamic parameters (73) (40), the actual magnitudes of the entropies and enthalpies of activation appear to be consistent with average values of the thermodynamic parameters associated with these mechanisms. Table XII gives some typical values of ΔS^\ddagger and ΔH^\ddagger for $B_{Ac}2$ and S_N1 type mechanisms.

Solvent Isotope Effect

The use of solvent isotope effects (SIE) as a

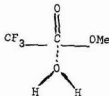
TABLE XII

SOME TYPICAL VALUES OF ΔS^\ddagger AND ΔH^\ddagger PARAMETERS FOR NEUTRAL HYDROLYSIS

Type	Compound	T °C	ΔH^\ddagger (Kcal. mole ⁻¹)	ΔS^\ddagger (Cal. mole ⁻¹ deg. ⁻¹)	Reference
S _N 1	tert.-butyl chloride	10	23.80	14.4	(70)
	tert.-pentyl chloride	10	21.98	10.2	(85)
	α -bromoisobutyrate ion	25	28.96	20.6	(73)
(?)	tert.-butyl dimethyl- sulphonium ion	50	31.76	15.7	(85)
B _{Ac} 2	chloromethyl chloro- acetate	25	11.8	- 37	(40)
	chloromethyl formate	25	11.4	- 39	(40)

tool for mechanistic diagnosis in solvolytic reactions has not been particularly rewarding. For instance, Robertson has stated that the SIE is a blunt tool for distinguishing between S_N2 and S_N1 mechanisms. In the case of the tri-fluoroacetates in the present work, the value of 1.2 measured for the tertiary ester would point to an $S_N1 - S_N2$ mechanism for the solvolysis of this compound. No attempt has been made in this work to distinguish between S_N1 and S_N2 mechanisms on the basis of SIE. The primary and secondary esters are again classified by the SIE in the same way as by the entropies and enthalpies of activation. The solvent isotope effect is considerably larger than that for the tertiary ester, in the range 3.4 - 3.8 which is supposedly typical of $B_{Ac}2$ processes (61).

To determine what would produce such a large SIE some information concerning the transition state and mechanism of the reaction is needed. The hydrogen bonding method of the estimation of the SIE, based upon zero point energies, seems to account for a reasonable proportion of the measured ratio but, nevertheless, this method originally proposed by Bunton and Shiner (61) seems oversimplified. They suggest the following transition state model for the $B_{Ac}2$ mechanism in the hydrolysis of CF_3COOMe :



and predict a solvent isotope effect of 2.7. The experimental value obtained by Bunton et al. in 40/60, V/V, water/dioxane is 1.8 (63). The ratio obtained for this compound in the present work is 3.5 in pure water. The lower value in the mixed solvent system is possibly a consequence of the breakdown of the water-water interactions due to the presence of dioxane (59). It is also an indication that some SIE values measured in solvent mixtures cannot always be considered in the same light, but must take into consideration the effect of the added solvent on SIE.

In Table IX the data shows that both the entropies and enthalpies of activation contribute to the SIE. However, it would be impossible at this time to determine why these quantities contribute in the way they do. In summary, the SIE substantiates the mechanistic classification based on enthalpies and entropies of activation. In the following section the SIE's are further discussed in relation to the heat capacities of activation for solvolysis reactions.

Heat Capacities of Activation

The heat capacity of the hydrolysis of ethyl trifluoroacetate has been calculated from the data in Table XII. The value of ΔC_p^\ddagger for this reaction is -66.99 cal./mol./deg/ which bisects the range (-30 to -120 cal./mol./deg.) obtained by Robertson (73) and other workers who have measured this parameter for a relatively wide range of solvolytic substrates. The magnitude of ΔC_p^\ddagger for this reaction is consistent with the idea that the solvolytic process involves the formation of a highly polar transition state from an initially neutral molecule and that the solvent reorganization brought about by the electric charges associated with the transition state is responsible for almost the entire value of ΔC_p^\ddagger (73), (86). Although the variation of the magnitude of this parameter with structural variations in the substrate is of interest, the limited heat capacity data allows only a qualitative assessment of the significance of this parameter.

From earlier evidence reported (20) the spontaneous hydrolysis of the ethyl trifluoroacetate takes place by a $B_{AC}2$ mechanism. From the experimental evidence in this work it appears that the reaction mechanism cannot be deduced from the ΔC_p^\ddagger value obtained, i.e. the heat capacity of activation

alone will not distinguish between the $S_N1 - S_N2$ and $B_{AC}2$ type mechanisms in a series of trifluoroacetates. Other parameters such as the SIE are apparently more useful for this purpose.

Smaller absolute values of ΔC_p^\ddagger than that for ethyl trifluoroacetate may be expected in cases where the solvation is not as great, e.g. in a reaction between a neutral molecule and an ion (86). In a second case small or positive ΔC_p^\ddagger values are possible when two mechanisms are occurring concurrently.

Kohnstam (86) has discussed this latter case briefly in a recent review. He has derived an expression for the observed heat capacity of activation using the Arrhenius expression for a reaction which occurs by a dual mechanism where the rate is $k = k_1 + k_2$. From his expression Kohnstam concludes that if the activation energies for the two reactions are not equal, then $\frac{k_1}{k}$ and $\frac{k_2}{k}$ will vary with temperature and consequently, so will $(\Delta C_p^\ddagger)_{obs}$. In most cases the change in $(\Delta C_p^\ddagger)_{obs}$ with temperature will be small. However, in the mechanistic borderline of nucleophilic substitution, where there is a possibility of two concurrent reactions of vastly different energies of activation, a temperature dependent heat capacity of activation may be expected.

A preliminary evaluation of the temperature dependence of the energy of activation for tertiary butyl trifluoroacetate indicates that the value of ΔC_p^\ddagger is positive. This fact is clearly shown by Figure 12. This leads logically to the assumption that the tertiary ester may be reacting by two, or possibly three mechanisms. The positive ΔC_p^\ddagger , as Kohnstam (86) suggests, leads to complications in the interpretation of the value of this parameter.

The effect of the two mechanisms on the heat capacity of activation can be treated in the following way.

Assuming that two mechanisms (rate constants k_1 and k_2) are involved, the effective rate constant is given by:

$$k_e = k_1 + k_2 \quad (35)$$

$$\frac{kT}{h} K_e^\ddagger = \frac{kT}{h} K_1^\ddagger + \frac{kT}{h} K_2^\ddagger \quad (36)$$

and hence

$$K_e^\ddagger = K_1^\ddagger + K_2^\ddagger$$

Assume that at $T = 25^\circ\text{C}$

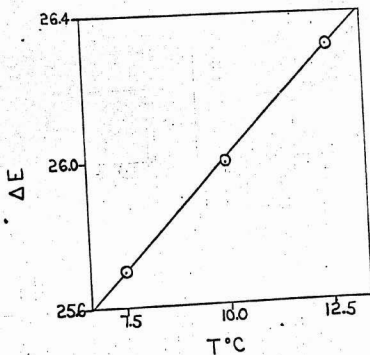
$$\Delta H_1^\ddagger = 11,000.0 \text{ cal/mole}, \Delta S_1^\ddagger = -35.00 \text{ e.u.}$$

which are typical B_{Ac}^2 parameters and

$$\Delta H_2^\ddagger = 25,000.0 \text{ cal/mole}, \Delta S_2^\ddagger = +15.00 \text{ e.u.}$$

Figure 12

CORRELATION OF THE ARRHENIUS ACTIVATION ENERGY VS
TEMPERATURE FOR $\text{CF}_3\text{COOC}(\text{CH}_3)_3$



also for the S_N1 case.

$$\Delta C_{p(1)}^{\dagger} = -65 \text{ cal./mole/deg., and}$$

$$\Delta C_{p(2)}^{\dagger} = -90 \text{ cal./mole/deg.}$$

The expression for the effective enthalpy of activation may be given by:

$$\Delta H_e^{\dagger} = \frac{K_1^{\dagger} \Delta H_1^{\dagger} + K_2^{\dagger} \Delta H_2^{\dagger}}{K_1^{\dagger} + K_2^{\dagger}} \quad (37)$$

$$\text{where } K_1^{\dagger} \cdot e^{-\Delta G_1^{\dagger}/RT} = e^{\Delta S_1^{\dagger}/R} e^{-\Delta H_1^{\dagger}/RT} \quad (38)$$

$$\text{and } K_2^{\dagger} = e^{-\Delta G_2^{\dagger}/RT} = e^{\Delta S_2^{\dagger}/R} e^{-\Delta H_2^{\dagger}/RT} \quad (39)$$

and the effective heat capacity is

$$\Delta C_{pe}^{\dagger} = \frac{K_1^{\dagger} K_2^{\dagger} (\Delta H_1^{\dagger} - \Delta H_2^{\dagger})^2}{RT^2 (K_1^{\dagger} + K_2^{\dagger})^2} + \frac{K_1^{\dagger} \Delta C_{p(1)}^{\dagger}}{K_1^{\dagger} + K_2^{\dagger}} + \frac{K_2^{\dagger} \Delta C_{p(2)}^{\dagger}}{K_1^{\dagger} + K_2^{\dagger}} \quad (40)$$

(a)
(b)
(c)

The expression for ΔC_{pe}^\ddagger may be evaluated in two ways: (i) Assuming that $\Delta C_p^\ddagger(1)$ and $\Delta C_p^\ddagger(2) = 0$, then the terms (b) and (c) are zero and the value of the measured heat capacity will be solely due to the difference between the enthalpies of activation. This non-genuine heat capacity is always positive and is caused by the simultaneous occurrence of the two mechanisms. (ii) On the other hand, if $\Delta C_p^\ddagger(1)$ and $\Delta C_p^\ddagger(2)$ are finite and negative and if

$$\frac{\partial \Delta C_p^\ddagger(1)}{\partial T} \quad \text{and} \quad \frac{\partial \Delta C_p^\ddagger(2)}{\partial T} = 0, \text{ the observed heat capacity}$$

will show peculiar behaviour and vary with temperature. Depending upon the magnitudes of the terms (a), (b), and (c) (equation 40) the value of ΔC_{pe}^\ddagger will change sign at some temperature. Whether the value of the parameter changes sign in regions that are experimentally accessible is not known and will depend on all three terms.

Queen (50) has found for the hydrolysis of propyl chloroformate that the heat capacity of activation is a positive quantity of 2.2 cal/mol./deg. and for ethyl and iso-propyl chloroformates that the ΔC_p^\ddagger values are close to zero although they are negative. It is possible, in these cases, that the reactions occur by concurrent mechanisms and that the available experimental range makes possible the detection of a change in sign of ΔC_{pe}^\ddagger .

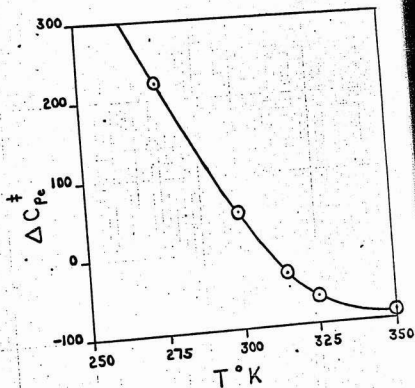
In the present theoretical case, depending on the size of the quantity $(\Delta H_1^\ddagger - \Delta H_2^\ddagger)^2$ the value of ΔC_{pe}^\ddagger will change and possibly remain negative but shift around with variations in structure in ways which partially depend on the relative importance of the two mechanisms at any temperature to the exclusion of solvation and other effects which only account for $\Delta C_p^\ddagger(1)$ and $\Delta C_p^\ddagger(2)$.

To show how ΔC_{pe}^\ddagger may change with T, calculations have been carried out using the values given earlier for the relevant parameters in Equation (40). Figure 13 shows that the effective heat capacity will approach the value for the higher heat capacity, i.e. the mechanism with the higher heat capacity will predominate at higher temperatures, and the lower heat capacity mechanism will become less important. However, depending on the region of accessible experimental measurements all three terms (a), (b), and (c) may be found to contribute extensively to the effective heat capacity of activation.

At present, it is not possible to determine precisely the contribution to the observed parameters from both mechanisms and to determine separate enthalpies of activation for both reactions. Care must therefore be taken in the interpretation of heat capacities in cases where dual mechanisms

Figure 13

GRAPHICAL REPRESENTATION OF ΔC_p FOR A REACTION
INVOLVING TWO SEPARATE PATHS



are suspected, irrespective of the type of mixed mechanism in question.

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